

REFORM PRACTICES IN CHEMISTRY EDUCATION
IN INTRODUCTORY COLLEGE AND
UNIVERSITY COURSES

By

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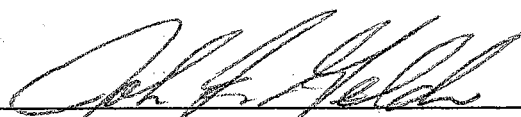
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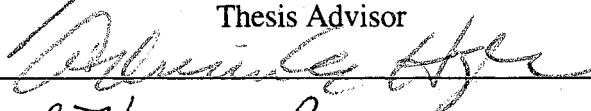
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CHAPTER 1

INTRODUCTION

To call for reform in science education is nothing new. In her study of the history of biology education, Dorothy Rosenthal cites an 1891 educator who bemoans a lack of emphasis on teaching students to think for themselves¹. In 1891, however, few science educators would have attempted to investigate experimentally which of several pedagogies was most effective. What distinguishes modern arguments for reform in science education is that the prescriptions are supported by a foundation of research in teaching and learning.

Like all disciplines, science education has experienced an information explosion in the second half of this century. Pedagogical models, and their philosophical underpinnings, are continually tested in an ever-increasing number of science education experiments. The keywords “science education” yielded 1336 records published in 1970 that are indexed in the Educational Resources Information Center (ERIC) database. When combined with the keywords “college, university or post-secondary,” the number of records was only reduced to 464. In 1990, there were 2684 new science education records indexed in ERIC, 1138 of which referred to higher education. Clearly, the number of research articles devoted to science education has grown.

It is difficult for the most ambitious, dedicated professors to keep pace with science education research, even that which is directly relevant to their discipline. For example, a search of ERIC records from 1970 reveals that thirty-three books and articles were published concerning post-secondary chemistry education. In 1990, the number of new publications listed on ERIC regarding higher education in chemistry had grown to over two hundred. At the same time, the body of chemical knowledge is increasing, perhaps even more rapidly. The amount of information which chemistry faculty “should” read is simply overwhelming.

It is difficult to quantify or compare student outcomes. Therefore, if instructional effectiveness is an important factor in tenure or evaluations, it is likely to be assessed in terms of perceived teaching effectiveness as measured by student surveys or peer opinion. This may remove the incentive for science faculty, especially those that are highly regarded by students and peers, to habitually peruse the science education research literature.

Fortunately, many professional societies regularly form committees to study and report on the state of science education. Professional societies’ articulations of the necessity of and appropriate directions for reform serve at least two purposes. First, they are a synthesis of consensus about education. The committee weighs conflicting evidence in the research literature and presents balanced conclusions and implications. Second, education reports from professional societies enhance dissemination of science education research

results because the publications in which they appear are much more widely read within the scientific community than education journals.

There is a large body of research in science education that suggests some pedagogical strategies are superior to others. This research has been condensed and interpreted in a variety of science journals. However, it is not clear how much this information has influenced the practice of teaching science at the college level. This dissertation has two major goals: 1) to report the extent to which existing course designs reflect professional societies' reform recommendations, and 2) to present a design for an introductory chemistry course which more fully addresses the objectives and goals of the reform recommendation than occurs in traditional curricula.

CHAPTER II

REVIEW OF THE LITERATURE

The first goal of this project is to determine whether educational research results as they are depicted by professional societies are changing the practice of science education in our colleges and universities. Before we can compare existing course designs and pedagogical strategies to the reform recommendations of professional societies, we must determine what those recommendations are. This chapter examines a variety of communications from scientific societies on science education, organizing the specific suggestions into broad themes. The reports summarized were published in the two decades from 1970 to 1990 by professional societies that represent a range of interests: general science, chemistry and engineering.

The second goal is to develop an example of an introductory chemistry course that fully reflects those reform recommendations. There is a rich body of information relevant to improving the quality of undergraduate chemistry education. This chapter also includes a review of the literature pertaining to fundamental changes in foundation courses as well as new approaches to pedagogy in chemistry at the post-secondary level.

Reform Recommendations for Science Courses

Appropriate goals for science courses have been suggested by a number of professional societies. Reports from the American Association for the Advancement of Science (AAAS), the American Chemical Society (ACS), the National Science Foundation (NSF), the National Science Teachers Association (NSTA), Sigma Xi, and the American Society for Engineering Education (ASEE) were examined and are paraphrased here.

A 1986 analysis of science education by AAAS concluded that practicing scientists need to be more involved in education and goals of math and science education must be clarified.² More detailed recommendations appeared in 1989.^{3,4} AAAS maintains that ideas and thinking skills must be emphasized at the expense of content details because quality of understanding is more important than quantity. Students should be active participants in science courses that begin with questions about nature and focus on collecting, evaluating and using evidence to answer those questions. AAAS stresses that “knowing should never be separated from finding out.”³ In addition, instructors must demand clear expression and classwork should model professional expectations, including the necessity for effective teamwork.

ACS issued guidelines for introductory chemistry courses in 1984.^{5, 6} Their principal recommendations were that investigative lab work should involve at least 30 % of class

time; curricula should emphasize properties and reactions of materials that students will encounter in everyday life; and courses should convey how hypotheses are formed, tested and eventually described as chemical “laws,” which includes an appreciation of the open-ended, falsifiable nature of chemical knowledge. These prescriptions were expanded in 1987 to include advocacy for research projects as early as the freshman year, a reexamination of the general chemistry curriculum, expansion and upgrading of labwork with access to modern instrumentation, and attention to group research and communication skills.⁷

In 1985, the NSF published two advisory statements on science education.^{8, 9} Co-op programs and summer internships to enhance skills and supplement lab instruction for students and faculty were a major focus of one report.⁸ The other article addressed broader issues in science education, recommending that colleges develop relationships with secondary and elementary schools, make early lab classes more attractive to students and provide more undergraduate research opportunities.⁹ In 1990, another analysis by the NSF suggested that laboratory instruction should be enhanced so that experiments interest students in science, develop important techniques and skills such as the ability to follow a written set of procedures and communicate the results, help students understand why we do experiments in science, and teach lab safety and proper disposal procedures.¹⁰ NSF also proposes that college instrumentation needs to be updated, under-represented groups must be attracted to and retained in science, school/industry interactions should be

encouraged, teaching and research need to be balanced and research on effective instruction should be expanded.

As early as 1970, the NSTA emphasized the need for students to understand the principles, processes, content and social implications of science.¹¹ NSTA also submitted that students should consider the usefulness of rational thought in making value-laden choices and be able to transform solutions into action. In 1981, recommendations included upgrading teacher education, designing applications-oriented curricula, recruiting under-represented groups into science, upgrading college equipment, developing instructional materials which incorporate computers, and conducting research on learning in science and math.¹² A lengthy 1987 NSTA analysis of science education states that students must be able to comprehend factual prose, organize information into useful structures, decide whether data support a statement, separate correlation from cause and effect and observation from inference, acquire information inductively and deductively, and use math, statistics and computers in problem solving.¹³ Students also need to practice collecting, interpreting, analyzing and evaluating data; making predictions from estimated data; generating many possible eventualities from a set of data; applying concepts and principles to new situations; and obtaining new information independently from a range of resources. They should be able to use scientific information within a multidisciplinary approach to make decisions, and participate as informed citizens with an awareness that real issues are

complex and precise outcomes are impossible to predict. This NSTA report also states that students should communicate effectively, exhibit a spirit of inquiry, work individually and as part of a group, exercise scientific values, and be able to distinguish between inference and reasoning from empirical evidence.

Sigma Xi examined science education in 1989¹⁴ and concluded that science courses need to be adjusted to students' preparation, hands-on investigations should be a part of every course, and more undergraduates should participate in research. Additional information published in 1990¹⁵ suggested that science education should encourage students to pursue science careers, and participate as informed citizens. The report insists that universities must focus on teaching undergraduates; rewarding, accessible curricula must be developed; classes should be smaller to encourage interaction and eliminate the "weed-out" mentality; students need more hands on experience with investigating phenomena; and curricula should be more flexible.

The ASEE focuses on the need for undergraduate research opportunities in a 1981 diagnosis of problems in engineering education.¹⁶ Recommendations in a 1986 report include teaching students to do experimental work of professional caliber, and developing courses which teach teamwork and communication skills.¹⁷ A 1987 opinion bemoans proliferation of courses and suggests that curricula should be reduced to the basics required for career-long professional learning.¹⁸

It is clear that the modifications envisioned range from sweeping structural changes in our educational system to minor adjustments in teaching style. Despite the unevenness of breadth and depth between and even within reports, several common themes emerge.

The most universally agreed upon reform is a shift from a factually dense curriculum to a more process-oriented or investigative curriculum. Recommendations acknowledge that making room for process skills in crowded courses will necessitate deleting some content. The first theme, transition to a more process-oriented curriculum, was articulated by every professional organization studied except for Sigma Xi. Specific support for this theme is detailed in Table I.

TABLE I

THEME I: SCIENCE COURSES SHOULD FOCUS MORE ON PROCESS

Recommending Body	Specific Recommendations
AAAS	<ul style="list-style-type: none"> - Goals of science education must be clarified.² - Ideas and thinking skills must be emphasized, at the expense of content details; quality of understanding is more important than quantity.^{3,4} - Collecting, evaluating and using evidence should be central to science courses.^{3,4} - Knowing should not be separated from finding out.^{3,4}

TABLE I (continued)

Recommending Body	Specific Recommendations
ACS	<ul style="list-style-type: none"> - Curricula, especially general chemistry, need to be reexamined.⁷ - Courses should convey how hypotheses are formed, tested and eventually described as chemical “laws.”^{5, 6} - Courses must communicate the open-ended, falsifiable nature of chemical knowledge and the experimental basis of knowledge.^{5, 6}
NSF	<ul style="list-style-type: none"> - Lab should teach students why scientists do experiments.¹⁰
NSTA	<ul style="list-style-type: none"> - Students must be able to comprehend factual prose, organize information into useful structures, decide whether data supports a statement, differentiate correlation from cause and effect and observation from inference, acquire information inductively and deductively, and use math, statistics and computers in problem solving.¹³ - Students need to practice making predictions from estimated data. They should understand that real issues are complex and precise outcomes are impossible to predict.¹³ - We should teach students to apply concepts and principles to new situations, and obtain new information independently from a broad range of resources.¹³ - Students should be able to distinguish between inference and reasoning from empirical evidence.¹³ - Courses should provide opportunities to collect, interpret, analyze and evaluate data.¹³
ASEE	<ul style="list-style-type: none"> - Curricula should be reduced to the basics required for life-long professional learning.¹⁸

The second most broadly endorsed idea is an increase in the quantity and quality of laboratory courses. This is closely related to Theme I, but the two do not completely overlap. A more investigative curriculum might include more laboratory, but an increase in the number of hours in laboratory does not necessarily mean that the experimental work is more investigative. Similarly, increasing the quality of laboratory sometimes implies more emphasis on process skills in concert with Theme I. However, sometimes increasing quality merely refers to updating equipment and instrumentation. Overall, recommendations focus on undergraduate research experiences, opportunities to use state-of-the-art equipment and investigative laboratory courses. Theme II, increased quantity and quality in laboratory instruction, was explicitly stated by all six professional societies as indicated in Table II.

TABLE II

THEME II: INCREASED QUANTITY AND QUALITY OF LABORATORY

Recommending Body	Specific Recommendations
AAAS	<ul style="list-style-type: none"> - Students should be actively involved and should begin with questions about nature.^{3,4} - Collecting, evaluating and using evidence should be central to science courses.^{3,4}

TABLE II (continued)

Recommending Body	Specific Recommendations
ACS	<ul style="list-style-type: none"> - Investigative lab work should involve at least 30% of class time in all introductory courses.^{5, 6} - Research projects should be included as early as the freshman year.⁷ - Lab work should be expanded and upgraded.⁷ - Students need better access to modern equipment.⁷
NSF	<ul style="list-style-type: none"> - Laboratory instruction needs enhancement to develop techniques and skills and teach lab safety and proper disposal; college instrumentation needs to be upgraded.¹⁰ - Co-op programs and summer internships are recommended for both students and faculty to enhance skills and supplement lab instruction.⁸ - Undergraduate research opportunities should be more available.⁸
NSTA	<ul style="list-style-type: none"> - Colleges need resources to upgrade equipment.¹¹ - Courses should provide opportunities to collect, interpret, analyze and evaluate data.¹³
Sigma Xi	<ul style="list-style-type: none"> - Hands-on investigations should be a part of every course.¹⁴ - More undergraduates should have the opportunity to participate in research.¹⁴ - Students need more hands-on experience with investigating phenomena.¹⁵
ASEE	<ul style="list-style-type: none"> - Opportunities for undergraduate research should be expanded.¹⁶ - Students should learn how to do experimental work of professional caliber.¹⁷

The third most commonly cited opportunity for improvement underscores the role of the educational institution as a professional training ground. Recommendations emphasize communication and group interaction skills that have not been traditionally considered the domain of science coursework. Theme III, development of skills students need as professionals, is supported by five of the six professional societies, all of which specifically cite communication skills.

TABLE III

THEME III: COURSES SHOULD DEVELOP SKILLS STUDENTS NEED AS PROFESSIONALS

Recommending Body	Specific Recommendations
AAAS	<ul style="list-style-type: none"> - Instructors must demand clear expression.^{3,4} - Professional scientists work as part of a team; classwork should model professional expectations.^{3, 4}
ACS	<ul style="list-style-type: none"> - Courses should develop research and communication skills.⁷
NSF	<ul style="list-style-type: none"> - Lab courses should teach students to follow a written set of procedures and communicate the results.¹⁵
NSTA	<ul style="list-style-type: none"> - Students need to be able to collect, interpret, analyze and evaluate data, both individually and as part of a group.¹³ - We must teach students to communicate effectively.¹³
ASEE	<ul style="list-style-type: none"> - Coursework should teach teamwork and communication skills.¹⁷

The fourth common thread we find relates to the number of scientists and engineers graduating in the United States. Professional societies periodically express concern that introductory sciences are perceived as “weed-out” courses, resulting in low science enrollments. Recommendations relate to making courses interesting, relevant or accessible. From Table IV we can see that no single suggestion is universally endorsed for Theme IV, that science courses should promote retention of students.

TABLE IV

THEME IV: SCIENCE COURSES SHOULD PROMOTE RETENTION

Recommending Body	Specific Recommendations
ACS	- Introductory curricula should emphasize properties and reactions of materials students will encounter in everyday life. ^{5, 6}
NSF	- Laboratory experiments, especially in introductory courses, should interest students in science. ^{9, 10} - Underrepresented groups must be attracted and retained . ¹⁰
NSTA	- Applications oriented curricula are needed for lower division college and high school courses. ¹² - More needs to be done to recruit under-represented groups into science. ¹²

TABLE IV (continued)

Recommending Body	Specific Recommendations
Sigma Xi	<ul style="list-style-type: none">- Science courses need to be adjusted to students' level of preparation; rewarding, accessible curricula must be developed.¹⁴- Science education should encourage students to pursue science careers, explore the nature and aesthetic qualities of science, engineering and math, and participate as informed citizens.¹⁴- Classes should be smaller to encourage student/teacher interaction; the 'weed-out' mentality must be eliminated.¹⁴- Curricula should be more flexible.¹⁴

The last recurring idea is Theme V, enhancement of teaching as a profession. Like the previous theme, suggestions diverge considerably for upgrading college and secondary education as professional activities. Table V illustrates the scope of recommendations.

TABLE V

THEME V: ENHANCING TEACHING AS A PROFESSION

Recommending Body	Specific Recommendations
AAAS	<ul style="list-style-type: none">- Teaching and research need to be balanced.^{3, 4}

TABLE V (continued)

Recommending Body	Specific Recommendations
<hr/>	
NSF	<ul style="list-style-type: none"> - Research on effective instruction should be expanded.¹⁰ - Colleges need to develop relationships with secondary and elementary schools.⁸
NSTA	<ul style="list-style-type: none"> - Teacher education needs upgrading.¹² - More research needs to be conducted on learning in math and science.¹³
Sigma Xi	<ul style="list-style-type: none"> - Universities must focus on teaching undergraduates.¹⁴

In summary, we can conclude that several measures are endorsed by professional societies that have studied the need for education reform in science. Increasing the quantity of laboratory experiences; focusing the factual content in courses and stressing process skills; and providing opportunities for students to improve communication skills are all supported extensively. Reform recommendations abound, and have been shown to have many commonalities, but there are no published studies or surveys to indicate whether any of these recommendations are actually being implemented.

During the last fifteen years there is abundant evidence of ferment about and interest in changes in chemistry curricula. Chemistry education publications can be divided into roughly two categories: articles which discuss entire degree programs or fundamental changes in foundation courses like general or organic chemistry, and articles which address particular courses. Although works in the second category are interesting, it is the first category with which we are primarily concerned.

A number of topics emerge which have garnered significant interest in the chemical education literature and which are relevant to entire chemistry programs rather than individual courses. Since 1980, there have been numerous articles on written communication skills (e.g., Rosenthal¹⁹, Melhado²⁰, Burkett²¹, and Wilson²²); considerable attention to whether and how much history of chemistry to include (e.g., Kamsar²³, Kauffman²⁴, and Zuckerman²⁵); many innovative suggestions for use of computers (e.g., Davis²⁶, Rasmussen²⁷, Johnson²⁸, Coleman²⁹ and Spain³⁰); extensive discussion on reviving descriptive chemistry, especially in the freshman course (e.g., Basalo³¹, Zuckerman³², Kenkel³³ and Ryan³⁴); recommendations on including more polymer chemistry in courses (e.g., Seymour³⁵, Mathias³⁶, Polymers Core Course Committee^{37,38}); a large body of articles on Piagetian approaches to pedagogy; and painful grappling with how to focus course content to cope with the information explosion (e.g.,

Wrighton³⁹, Kreyenbuhl⁴⁰, and Hawkes⁴¹).

All of these topics are relevant to general chemistry, but to try to focus on all of them simultaneously in one course is overwhelming. New approaches to pedagogy show promise as one of the most transforming of the general directions in curricular reform, therefore a review of the literature pertaining to this topic follows.

Many articles on writing across the curriculum propose that communication is not only essential to professionals, but is also a good way to evaluate and encourage higher level thinking skills, according to Bloom's taxonomy.⁴² The other educational strategy related to Bloom's work, which encourages mastery learning, is the Keller Plan, or Personalized Instruction.²⁶ Most other research in chemistry pedagogy depends on Piaget as the theoretical foundation.

The application of Piaget's work to chemistry education begins with the recognition of concrete and formal operational stages of reasoning. Since much of chemistry involves manipulating abstractions, students at the concrete operational level can be expected to have difficulty with the material.⁴³ In addition, the notion that we re-enter concrete operational thought when on unfamiliar intellectual territory helps explain why chemistry is difficult for such a large proportion of students.⁴⁴ According to Bodner, "Piaget argued that knowledge is constructed as the learner strives to organize his or her experiences in terms of pre-existing mental structures or schemes."⁴⁵ The traditional view of learning is that

knowledge is transmitted from teacher to learner. A Piagetian view that knowledge is constructed in the mind of the learner and is based on the learner's previous experience suggests that many traditional teaching methods, such as the lecture, may not be extraordinarily effective at promoting learning.⁴⁶

Active learning strategies that have been developed as alternatives to traditional lecture include group assignments both in and out of class,^{47,48} discussion based on pre-class or in-class assignments,^{49,50} in-class writing,⁴⁷ lectures interspersed with "breaks" for brief assignments completed by pairs or small groups,⁵¹ learning cycle approaches that emphasize an exploratory stage in laboratory,^{34,52} extra- or co-curricular undergraduate research,⁵³ integrated curricula that encourage students to establish connections between the subject matter and relevant experience,^{54, 55} material presentation structured around "marathon problems,"⁵⁶ and laboratory experiments that are more investigative or less "cookbook."^{57, 58}

Despite the educational research supporting the use of active learning pedagogy, including a study showing that students who attended lectures did no better on exams than students who were absent,⁵⁹ the traditional lecture has by no means been discredited or eliminated. In chemistry education research, true experimental controls are elusive. This makes comparing the efficacy of standard lecture to an alternative difficult, and means that comparing different alternatives to each other is nearly impossible. In addition, instructors

agree that adopting Piagetian teaching strategies means a reduction in the amount of content that can be presented in a course.⁶⁰

Chemistry teachers do not have unambiguous proof that alternatives are more effective than lecture, and they have even less information about which alternative is the best. Unlike some secondary schools, colleges and universities do not have statewide curricula nor centralized textbook choice. Hence, university curricular change is not driven by university administration, but rather by chemistry departments and individual faculty. The only administrative structure with significant influence over chemistry curricula is the American Chemical Society, which certifies chemistry departments. This influence is limited however, because the certification standards have been mostly concerned with resources available, such as number of faculty and range of courses offered, rather than pedagogy or curricular details within a given course. No professional society's recommendations have ever contained a detailed course outline; reform recommendations are general and assume the sacrosanctity of academic freedom. As a result, instructors who do adopt active learning strategies generally use a mix that they feel will maximize both content coverage and student involvement in learning.

CHAPTER 3

METHODS

Have educational reform recommendations had any impact on college and university science programs? Some understanding of the extent to which science curricula have integrated suggested reforms can be obtained by examining course syllabi. This chapter describes how syllabi were gathered and analyzed for implementation of innovations suggested by scientific societies. A rationale for and method of construction of a new introductory chemistry curriculum is also provided.

Investigation of Reform Implementation in Science Courses

The question under examination is whether college and university science programs are responding to the concerns of their professional societies. Course syllabi will furnish some insight into the answer to this question. Statements on a syllabus may not be achieved, but they should represent the instructor's intentions for the course. A review of syllabi for one course from a large number of colleges and universities provides a manageable means of

investigating the extent of curricular reform in science courses.

The colleges and universities which appeared in the first quartile of *U.S. News and World Report's* 1993 ranking⁶¹ of undergraduate colleges and universities were solicited between May of 1994 and April of 1995 for a copy of their general chemistry course syllabi. The *U.S. News* ranking was used to determine the sample group because it was the most recently available ranking and is based on a broad range of criteria. Since these first-quartile institutions have strong reputations and solid resources, one would expect more innovative curricula than at less prestigious institutions with more limited resources. To verify that this assumption was not completely erroneous, a smaller sample of syllabi were solicited from schools identified as third-quartile by the 1993 *U.S. News* ranking.

Syllabi for an introductory chemistry course with laboratory were requested by course number. Appropriate course numbers for each institution were identified by examining college catalogs to ensure that syllabi received would be comparable. The course number chosen represented that institution's introductory chemistry class for chemistry majors.

As noted in Chapter Two, three reform recommendations received support from a broad range of sources. Syllabi were examined for evidence of implementation of one of these three reform recommendations: 1) increasing the quantity of laboratory experiences; 2) focusing the factual content in courses and stressing process skills; or 3) providing opportunities for students to improve communication skills.

Syllabi were scored as either providing no evidence of reform efforts, or as providing evidence for attempting to implement some type of reform. Syllabi which were scored as demonstrating reform contained one of the following measures that correspond to the reform recommendations listed above:

- 1) increased time in laboratory, with one three-hour period per week as standard,
- 2) increased quality of laboratory experiences, defined as more investigative or less “cookbook” than most commercial laboratory manuals’ experiments, or
- 3) writing assignments or oral presentations.

These measures do not ensure that any particular outcome will be achieved. They are only used in this study as indicators of an active attempt to respond to professional societies’ perception of the need for change in science education.

Four categories of institutions were solicited for syllabi: fifty-one first-quartile universities, thirty-five first-quartile colleges, seventeen third-quartile universities, and twelve third-quartile colleges. After two requests by mail, one of the four categories (third-quartile colleges) had less than a 33% return rate. Subsequent contacts to third-quartile colleges were made by telephone until one-third had provided complete information. A small sample of institutions who did not return complete syllabi were interviewed by telephone to ascertain whether there was likely to be a self-selection bias in the institutions which did return complete information.

It was apparent from syllabi that a number of institutions use in-house laboratory manuals instead of commercial texts. Those institutions received a request by mail for a copy of their lab manual. Between February and April of 1995 institutions with in-house lab manuals that had not returned information about their laboratory program were contacted by phone. The laboratory coordinator or an appropriate faculty member was interviewed to determine whether the laboratory curriculum was more extensive or investigative than a traditional curriculum.

Developing Another Model for College Chemistry Curricula

Some nontraditional curricula, such as the Holy Cross Discovery Chemistry program,⁶² are available, but the number of models for general chemistry programs are limited. In an effort to provide one more option that is not a standard lecture plus lab course, the information in the first semester of chemistry is presented in a way that responds to some of the reform recommendations that receive broad support from a range of professional societies. The basic content and order are the same as a traditional course, but topics are introduced through lab, assignments are more attentive to science process skills and communication skills are intentionally taught.

Some of the laboratory assignments are based on classic experiments in general

chemistry. However, instead of step-by-step instructions, students are given some background information and asked to either formulate and test a hypothesis or to collect some data and then generate a hypothesis from that data. Where classical experiments could not be easily modified for this format, new experiments were designed.

After laboratory, the students' data and hypotheses are the starting point for subsequent lectures or discussions. According to Piaget's model of learning, this concrete background should enhance the majority of students' understanding relative to introducing a topic in the abstract setting of lecture. Shifting the focus to laboratory also provides opportunity for students to practice skills crucial to real science, including collecting, evaluating and using evidence. Traditional curricula include laboratory but the assignments are often limited to collecting data to verify empirical relationships that have already been defined in the lecture portion of the course. Here, the "hypothesis" is not generated by the student, but is already a given. In contrast, in real science, preliminary data is often used to generate hypotheses, which are then tested more extensively for verification or modification. By introducing topics in lab and by designing introductory labs to be open-ended or investigative in nature, students have ample opportunities to generate and test their own hypotheses and to evaluate the quality of their data. This strategy not only teaches important thinking skills; from a Piagetian point of view, it is superior to traditional laboratories. Allowing students more independence in lab forces them to become more intellectually involved, so they will

probably better retain what they learn and do.

Ultimately, scientists communicate by publishing in journals and presenting papers at conferences. Using professional notebook standards is the first step toward that horizon.

Although grading notebooks is time-consuming, report sheets are generally inadequate for open-ended laboratory assignments anyway. A requirement was also included for students to present the results of an experiment to the class in a mini-symposium format.

CHAPTER 4

RESULTS

This chapter presents the results of the survey of syllabi from colleges and universities. First-quartile colleges and universities are used as the sample group, assuming that other categories of institutions will not exceed their accomplishment in terms of implementing reforms. Smaller samples of third-quartile colleges and universities are examined to test this assumption. A telephone survey of nonrespondents is used to check for self-selection bias of respondents.

This chapter also provides a description of an alternative general chemistry curriculum. A detailed outline of notes and laboratory instructions for this first-semester course is included as Appendix A.

Investigation of Reform Implementation in Science Courses

Table VI lists first-quartile universities from which complete information was received, in the order that they appear in the 1993 *U.S. News* ranking.⁶¹ All first-quartile

universities received a request by mail for syllabi from their general chemistry course.

Seventeen of the fifty-one institutions returned complete information: either a lecture syllabus that included laboratory information, or a lecture syllabus plus a laboratory syllabus. Table VI denotes whether a phone contact was made. It also tabulates whether an in-house laboratory manual was indicated on the syllabus and/or received.

TABLE VI
INFORMATION RECEIVED FROM FIRST-QUARTILE UNIVERSITIES

Institution	Lecture Syllabus Received	Lab Syllabus Received	In-House Lab Manual	Phone Contact
Duke University	Includes Lab Information		X	X
Cornell University	Includes Lab Information		X	X
Columbia University	X	X	X	X
Rice University	X	X	X	X
Univ. of Pennsylvania	X	X		
Washington University	X	X	X	X
U. Cal. at Los Angeles	X	X	Received	
Emory University	Includes Lab Information		Received	
Boston University	X	X		

TABLE VI (continued)

Institution	Lecture Syllabus Received	Lab Syllabus Received	In-House Lab Manual	Phone Contact
Coll. William and Mary	X	X	X	X
SUNY at Binghamtom	Includes Lab Information		X	X
U. Cal. at San Diego	X	X	X	X
Univ. of Cal. at Irvine	X	X	Received	
U Ill. Urbana-Champaign	X	X	Received	
U. N. C. Chapel Hill	X	X	Received	
Univ. of Washington	X	X	X	X
U. Cal. Santa Barbara	Includes Lab Information		Received	

Thirty-five liberal arts colleges are listed as first-quartile by *U.S. News*.⁶¹ Of these, fourteen responded to the survey with complete data. Table VII details the institutions which returned complete information, in the order that they appear in the 1993 *U.S. News* ranking, and the type of data received.

TABLE VII
INFORMATION RECEIVED FROM FIRST-QUARTILE COLLEGES

Institution	Lecture Syllabus Received	Lab Syllabus Received	In-House Lab Manual	Phone Contact
Williams College	Includes Lab Information		X	X
Swarthmore	Includes Lab Information		X	X
Smith College	Includes Lab Information		X	X
Davidson College	Includes Lab Information		X	X
Grinnell College	Includes Lab Information		Received	
Oberlin College	X	X	Description Rcvd.	
Hamilton College	Includes Lab Information		Received	X
Trinity College	Includes Lab Information		Received	
Coll. of the Holy Cross			Received	
Barnard College	X	X	X	X
Colorado College	Includes Lab Information		X	X
Connecticut College	Includes Lab Information			
Macalester College	Includes Lab Information			
Union College	Includes Lab Information		Received	X

Seventeen third-quartile universities, which is one-third of the fifty-one institutions listed by *U.S. News*, received a request by mail for syllabi. Eight universities returned complete data. Institutions in this category were solicited to rule out the possibility that first-quartile universities, even with superior reputations and resources, are actually less innovative than institutions that are not widely acclaimed for excellence. A smaller sample size was adequate for this between-group comparison.

TABLE VIII
INFORMATION RECEIVED FROM THIRD-QUARTILE UNIVERSITIES

Institution	Lecture Syllabus Received	Lab Syllabus Received	In-House Lab Manual	Phone Contact
Arizona State University	Includes Lab Information			
Drake University	Includes Lab Information		X	X
Florida Institute Tech.	Includes Lab Information			
Kansas State University	Includes Lab Information		X	X
Oregon State University	Includes Lab Information		X	X
Univ. Illinois at Chicago	Includes Lab Information			
Univ. of North Dakota	X	X		
Washington State Univ.	Includes Lab Information		X	X

Twelve colleges, or one-third of the thirty-five ranked as third-quartile, received requests by mail for general chemistry syllabi. After two mail requests, only one institution had returned complete information. Subsequent contact was made by phone. In every case, one of the faculty currently teaching general chemistry was interviewed. If that faculty member was not responsible for laboratory, the lab co-ordinator was interviewed as well. Table IX denotes which institutions were successfully contacted by telephone.

TABLE IX
INFORMATION RECEIVED FROM THIRD-QUARTILE COLLEGES

Institution	Lecture Syllabus Received	Lab Syllabus Received	In-House Lab Manual	Phone Contact
Austin College				X
Chatham College				X
Juniata			X	X
Oglethorpe University	Includes Lab Information			

Table X itemizes the proportion of institutions within each category from which complete information was obtained. It also notes the proportion of institutions which were

contacted by telephone. All four categories have response rates between one-third and one-half. A similar proportion of first-quartile universities, first-quartile colleges and third-quartile universities (between 50 and 57 percent) was contacted by phone to determine the content of in-house laboratory manuals. A higher proportion of third quartile colleges was contacted by phone to collect information obtained from syllabi for other institutional categories.

TABLE X
PROPORTION OF RESPONDENTS CONTACTED BY PHONE

Institution Category	Number Solicited	Number of Complete Responses	Number of Complete Responses Contacted by Phone	Percent Complete Responses	Percent Phone Contacts
First Quartile Universities	51	17	9	33	53
Third Quartile Universities	17	8	4	47	50
First Quartile Colleges	35	14	8	40	57
Third Quartile Colleges	12	4	3	33	75

Syllabi were examined for evidence of one or more of the following:

- 1) increased time in laboratory, with one three-hour period per week as standard;
- 2) increased quality of laboratory experiences, defined as more investigative or less “cookbook” than most commercial laboratory manuals’ experiments; or
- 3) writing assignments or oral presentations.

No institution provided evidence that strategy 1 was implemented at the introductory chemistry level. A number of institutions included syllabi statements or lab activities from in-house laboratory manuals that addressed strategy 2. The implementation schemes varied, but could be assigned in each case to one of three general categories:

- 2a) Addition of one or two investigative laboratories to a fairly standard laboratory curriculum.
- 2b) Adoption of a series of investigative laboratories in place of a standard curriculum.
- 2c) Inclusion of a capstone project of an investigative nature in the introductory laboratory.

Several institutions employed strategy three by including writing assignments in lecture syllabi. It was often impossible to determine whether laboratory assignments were formal written reports or fill-in-the-blank report sheets, therefore, writing assignments in laboratory were not considered. Of the twenty-five universities which returned complete information, twelve showed some evidence of reform implementation. The specific type of curricular reform employed by each university is listed in Table XI.

TABLE XI
RESULTS OF SYLLABI ANALYSIS FOR UNIVERSITIES

Institution	One or Two Investigative Labs (Strategy 2a)	Series of Investigative Labs (Strategy 2b)	Capstone Project (Strategy 2c)	Writing Assignments (Strategy 3)
First Quartile Universities N=17				
Boston University		X	X	
Emory University	X			
SUNY-Binghamton	X			
U. California at Irvine	X			
U. California-San Diego				Honors Section
UI Urbana-Champaign	X			
Univ. of Pennsylvania	X		optional	
Univ. of Washington				X
Third Quartile Universities N=8				
Drake University	X			X
Kansas State University		Honors Section		
Oregon State University			X	
Washington State Univ.	X			

The majority of universities in Table XI include one or two investigative labs in an otherwise traditional laboratory curriculum. Eight of the ten universities that were attempting to enhance the quality of laboratory experiences also employed in-house lab manuals. The two exceptions were Boston University and the University of Pennsylvania. Boston University has adopted *Inquiries Into Chemistry, 2ed*,⁶³ which is one of the very few commercial laboratory manuals with optional investigative activities. University of Pennsylvania uses Chemtrek⁶⁴. Chemtrek is a small-scale laboratory manual, which means, according to the lab syllabus, that “‘tinkering’ is encouraged [therefore lab] is less ‘cookbook.’”

Table XII details the types of curricular reform found in first- and third-quartile colleges. Of the eighteen colleges which returned complete information, ten supplied syllabi indicating an attempt to respond to professional societies’ critiques. As in Table XI, the nontraditional aspects of the courses can be assigned to either strategy 2 or strategy 3. In contrast to universities, far more colleges report using a series of investigative labs, rather than just one or two investigative labs in an otherwise traditional laboratory program. Committing more of the total class time to investigative laboratories seems to indicate a more radical change in the laboratory curriculum. However, there appears to be less variety in colleges than universities in the kinds of approaches that are being used to transform science education. No capstone projects or writing assignments were apparent

from college syllabi, although both schemes were used by some universities. All ten of the institutions listed in Table XII have developed an in-house laboratory manual.

TABLE XII
RESULTS OF SYLLABI ANALYSIS FOR COLLEGES

Institution	One or Two Investigative Labs (Strategy 2a)	Series of Investigative Labs (Strategy 2b)	Capstone Project (Strategy 2c)	Writing Assignments (Strategy 3)
First Quartile Colleges N=14				
Coll. of the Holy Cross		X		
Colorado College		X		
Grinnell	X			
Hamilton College	1st semester	2nd semester		
Oberlin College		X		
Smith College		X		
Trinity College		X		
Union College	X			
Williams College		X		
Third-Quartile College N=4				
Juniata College		X		

In both colleges and universities, most attempts at curricular reform revolved around strategy two, enhancement of quality in undergraduate laboratories (see Tables XI and XII). All but two of the colleges and universities which are implementing nontraditional methods for teaching laboratory also employ an in-house laboratory manual. Universities' syllabi reveal a larger diversity of approaches, whereas colleges' laboratory manuals indicate a larger investment of class time in nontraditional structures.

Table XIII compares by institutional category the amount of curricular reform visible from syllabi. Frequencies of observed reform were compared between categories with a Chi-square test of multinomial probability distribution. As expected, at a ninety-five percent confidence level, third-quartile colleges and universities are not more likely than first-quartile colleges and universities to demonstrate evidence of reform. The difference in the amount of observed reform in first-quartile universities and first-quartile colleges is not significant at a ninety-five percent confidence level. However, as noted above, the approaches used in the two categories are not necessarily comparable.

TABLE XIII

**PROPORTION OF RESPONDENTS IMPLEMENTING
CURRICULAR REFORM**

Institution Category	Number of Complete Responses	Number of Complete Responses that Provide Evidence of Curricular Reform	Percent of Complete Responses that Provide Evidence of Curricular Reform
First-Quartile Universities	17	8	47
Third-Quartile Universities	8	4	50
First-Quartile Colleges	14	9	64
Third-Quartile Colleges	4	1	25

To determine whether the samples were representative, phone interviews were conducted with representative from seven institutions who did not respond to the request for syllabi. Only one reported any activity in general chemistry that would have been scored in this study. Therefore, it is likely that a self-selection bias exists in the sample and that actual reform implementation is not as widespread as the study would indicate.

College Chemistry Curricula

The following is a description of a first-semester general chemistry course that incorporates broadly supported reform recommendations. Included for reference is the table of contents (Table XIV) for the actual classroom notes and laboratory activities which appear in Appendix A. Features to note include:

- 1) Introduction of topics in laboratory. Generally new concepts are used in laboratory before abstract applications are discussed in class.
- 2) Student design of laboratory activities. Directions for laboratory are often general and ask students to design their own specific plan for accomplishing the stated goals.
- 3) Student use of professional resources. Part of the design of an experiment includes looking up safety and disposal information. Students learn to make choices about starting materials based partly on handling considerations.
- 4) Development of professional notebook standards. Student design of labs allows notebook standards appropriate to a research environment to be taught.

The topics covered and their order are quite traditional. The main distinction of this curriculum is the integration between lecture/discussion and lab and the open-ended, investigative nature of the laboratory component, both of which respond to reform recommendations from the literature.

TABLE XIV
TABLE OF CONTENTS
First Semester General Chemistry Course

Chapter	Appendix A, Page
I. Elements, Compounds and Mixtures	69
IA. Naming Compounds	72
IB. Safety and Disposal	74
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III. Measurements and Data	78
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XI. Properties of Gases	98
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XXI. Trends in the Periodic Table	122
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XXVI. Properties of Solutions	133
XXVII. Reactions Which Transfer Electrons	136

Section I, Elements, Compounds and Mixtures poses the question “What is an atom?”

Most students are able to provide a definition, because they have had some exposure to chemistry before. The typical response is “the smallest particle of something.” This does not distinguish between atoms and molecules, however. Students are then provided with a set of labeled jars. Each jar contains either an element, a compound or a mixture. Students record observations from examining, for instance, a jar of zinc metal, a jar of sulfur (a yellow powder), a mixture of zinc metal and sulfur, and a jar of zinc sulfide (a white powder). Students write definitions for element, compound and mixture based on their observations and compare their definitions to the textbook’s. The concrete reference of lab makes nuances in the textbook definition clearer. With a firm foundation in the difference between an element, a compound and a mixture, the terms atom and molecule can be addressed more meaningfully in lecture/discussion. Perhaps more importantly, students gain a sense of how early chemists might have developed these definitions and are on their way to understanding chemistry as an empirical science rather than a compendium of facts. This characteristic of the exercise makes it equally valuable for experienced and inexperienced chemistry students.

Section II, Atomic Masses asks “How much will a pound of magnesium weigh after oxidation to magnesium oxide?” The instructor explains that an oxygen atom is being

linked to every magnesium atom, so that the product will have twice as many atoms. The most common suggestion is that the MgO will weigh twice as much as the Mg. Regardless of the actual guess proposed, we then check our guesses by experiment.

Students measure the ratio of masses of two elements, magnesium and oxygen. Before we perform the experiment, we start with Section III, Measurements and Data, which is a short unit on significant digits, precision and accuracy, and chemical hazard labeling systems. Students are then asked to design their own procedures for determining the mass ratio of magnesium and oxygen. They are provided with the following information:

Experimental overview:

Many elements will combine with oxygen from the air when heated. Magnesium, a solid metal, will combine with oxygen to produce magnesium oxide (MgO) and with nitrogen to form magnesium nitride (Mg₃N₂). Water will convert magnesium nitride to magnesium hydroxide, releasing ammonia (NH₃) into the air. Heating magnesium hydroxide will form magnesium oxide, a white or gray solid. We will use these reactions to compare the mass of a magnesium atom with the mass of an oxygen atom, since MgO has equal numbers of each.

Together with your lab partner, design a method to fully convert magnesium to magnesium oxide. You will need to determine the final weight of the oxygen used in the compound, and the ratio of Mg mass to O mass. Make your procedures as detailed as possible and be sure to include hazard and disposal information for chemicals. Triple space your procedures so that you can add information.

Introductory material in Sections IA and IB provides the information students need to plan for calibrating balances, recording data correctly, and determining hazard and disposal

information before the experiment. The instructor provides a checklist of items that are essential, such as recording the mass of weighing paper. The checklist fulfills the purpose of step-by-step instructions: time is not wasted because a simple error makes results impossible to interpret. The students, however, generally understand their work since they have planned it, which in turn makes it easier for them to understand and interpret results. At the conclusion of the experiment, the instructor compiles the class data and presents it in lecture/discussion. The class average mass ratio is used to answer our initial question.

The point that an atom's mass depends on its identity is obvious to students. The masses of oxygen and magnesium are different enough that magnesium clearly is heavier. This provides conceptual background for a more detailed discussion of the periodic table, isotopes and the carbon-12 basis for the atomic mass unit. Students are asked to compare their reported values of the Mg/O mass ratio to the value determined from the periodic table. If the comparison is unfavorable, they have the opportunity to assess their procedures and suggest improvements.

The next question is "How many atoms of Mg did we start with in the last experiment?" or alternately, "How many atoms are in a mole?" This question is addressed in section V, Chemical Moles. Since we use unit analysis to teach stoichiometry, and since they need a good grasp of the metric system, our discussion of the mole is preceded by section IV, Metric Units. This segment on metric conversions also introduces unit analysis. The concept of the mole is explained in lecture/discussion, building on the already familiar fact

that different elements have different masses. Students are then asked to suggest a means for measuring the number of atoms in a mole. This is set up as a “thought experiment” and lab groups can use resources in their plans that are not actually available to them. Students present their experiments to their lab class which compares and critiques the experiments. The instructor points out whenever the suggested methods are similar or related to methods that have historically been used to measure the mole. Having a variety of experimental means to measure the mole described invariably makes the concept clearer. Students are then referred back to their magnesium oxide experiment and asked to determine the number of moles and molecules in their sample.

The next question we ask students is “How do we know the formula of magnesium oxide or water?” After a brief introduction which explains the use of subscripts in chemical formulas, the following information is provided in Section VI, Chemical Formulas.

Experimental Overview:

We used the mass of a reactant (magnesium) and a product (magnesium oxide) to determine the mass of the other reactant (oxygen). To accomplish this we had to know the formula of magnesium oxide (MgO , one magnesium for every oxygen). This time you will use masses of the reactants to determine the formula of the product. You will prepare a new compound by mixing together two solutions. Then you will determine how much of each substance is in the new compound.

A solution of iron will be provided, labeled with the number of grams of iron per mL of solution. Sodium hydroxide (a solid) is soluble in water. When iron and hydroxide solutions are mixed, a precipitate (solid) will form.

With your lab partner, develop a detailed set of procedures to determine the mass of each substance in your final compound. Make sure you consider how you will know whether all of your reactants were used to form the solid. Include chemical hazard and disposal information for any reagents you plan to use. Have the instructor check your completed procedures.

Express the amount of each substance in the compound as a percent of the total mass. Determine a formula from the masses you measured. Turn in your data.

This is a challenging exercise for students, but with a lab partner and some work they can usually figure out that the sodium hydroxide can be weighed before it is dissolved, that the number of grams of iron used is calculated from the volume of iron solution, and that both can be converted to moles. Since the iron solution is colored, students who have an excess of iron invariably figure this out for themselves while filtering their precipitate. The instructor questions lab pairs who have an excess of sodium hydroxide to make sure that they realize this.

Many students realize that they can convert their masses to moles and hence determine a formula. Some students only turn in a percent composition. The instructor presents the aggregate data to the class and uses it to determine a formula for iron hydroxide. Even the students who were not able to do this on their own are able to follow the instructor's reasoning much more readily since the lab experience reduces the abstraction.

The goal of Section VII, Chemical Reactions is to introduce the first law of thermodynamics, balancing of chemical equations, and classifying reactions and predicting

products. Conservation of mass and energy is presented in lecture/discussion. Four categories of reactions (acid-base: NaOH and HCl, using universal indicator to compare products and reactants; double displacement: cobalt (II) chloride and sodium phosphate, which produces a precipitate that is colored, allowing us to identify the precipitate; single displacement: Mg and HCl; and combustion: ethanol and oxygen) are presented as lecture demonstrations. Decomposition of calcium carbonate is also discussed. The instructor writes and balances an equation for each reaction. Students are then given the opportunity to practice in laboratory where a number of acid-base, double displacement and single displacement reactions have been set up at stations around the lab. Students rotate through the stations, trying the reaction, identifying the category and writing a balanced equation by using the examples in lecture as a rubric. The investigative portion of this laboratory revolves around determining the identity of the precipitate in double displacement reactions.

Sections VIII through X, Calculations in Chemistry, Using Stoichiometry, and Reactions in Solution are essentially one unit that introduces stoichiometric calculations. Students often find stoichiometry to be frustrating busywork and fail to see any real-world application for it, despite textbooks' laudable efforts to include relevant, applications-oriented review exercises. Undergraduates tend to learn stoichiometry by rote as a set of algorithms, rather than learning to reason chemically. The laboratory exercise in this unit is a direct application of stoichiometry as well as observation to answer a fairly interesting question: "Can baking soda degenerate into harmful products if it is heated too much?"

The iron hydroxide precipitation laboratory has already demonstrated the concept of limiting reactant. Students are asked to look back at their data and bring to class the mass of each reactant from their experiment and whether it was used up. Lecture/discussion provides an everyday analogy for limiting reactant problems and the instructor works several as examples. Students then calculate the theoretical yield of their iron (III) hydroxide and compare it to the actual yield.

With this background, students return to the laboratory. They are asked to bring to class a number of possible (balanced) equations to describe the decomposition of sodium bicarbonate. These are listed on the board and students are given the following instructions.

Experimental overview:

A list of possible reactions for the decomposition of NaHCO_3 is available. You should plan a strategy to help you determine which, if any, of these equations best describe decomposition of NaHCO_3 .

When you have a strategy recorded, discuss it with a neighboring lab group and then write out detailed procedures. You may perform a qualitative investigation before you plan detailed procedures. Since you do not know what your products are, save the products for disposal after they have been identified.

Students then plan how they will experimentally determine which, if any, of the possible equations is correct. Both the textbook and the CRC are references for investigating the properties of the potential products of the reaction. The instructor asks people to share

progress on their experimental plans and at least some students will suggest comparing the theoretical yield of the sodium-containing product to the actual mass of remaining solid. Students complete their plans with input from the class discussion, the instructor checks individual procedures for safety and hands out a checklist of reminders for experimental work, and students proceed to the actual experiment.

Lab is followed by discussion of the potential equations for the decomposition of baking soda and lecture presentation of typical stoichiometry problems and solution strategies. Various units of concentration are then defined and stoichiometric calculations are extended to include problems involving solutions of reactants and products.

In Section XI, Properties of Gases and Section XII, Kinetic Molecular Theory of Gas Behavior, we begin by asking students what they already know about gas behavior. Students reason from their own experience with everyday objects about relationships between gas temperature, pressure and volume. The instructor puts these relationships together into the ideal gas law and explains the ideal gas constant and its units. The next question is “What is the value of the ideal gas constant?” Students are sent into the laboratory with the following information.

Experimental Overview:

We will determine the value of the ideal gas constant, R . In groups of four, devise an experiment that will allow you to measure the quantities necessary to determine the ideal gas constant. When you have outlined a set of procedures, consult the troubleshooting checklist to correct any major omissions, and then have the instructor check your final procedures before

you begin work.

Groups of four perform this experiment so that the instructor can consult with each group, if necessary, as they formulate a plan.

Follow-up in lecture/discussion centers first around ideal gas equation problems and then addresses partial pressure example problems. At least a few students in each class will have designed ideal gas constant experiments that involved collecting a gas over water. The student and instructor explain how the experiment was conducted and present the data. The class then corrects that data in small groups in class or as homework, and the instructor goes over the corrected data in the next class period. The assumptions of KMT and rationalizations based on KMT for observed ideal gas behavior are then presented in lecture/discussion. Finally, conditions under which the ideal gas equation will not hold are reviewed.

Our next question is “Why do some reactions occur spontaneously, while others do not?” Students develop an answer using material in Sections XIII through XVI, Energy in Chemical Reactions, Measuring Enthalpy, Calculating Energy Changes, and States of Matter. Students are sent into the laboratory after brief introductory remarks on the relationship between temperature and heat, with the following directions.

Experimental Overview:

A number of reactants have been prepared for you. Record as many observations as possible, including the initial temperature. Mix the reactants together and record your observations. Write a balanced equation for the

reaction and determine whether the reactants gained or lost heat as they were converted to products. Write your initial and final temperature on the overhead next to the appropriate reaction.

One of the “reactions” is an endothermic solution process; the others are exothermic reactions. Lecture/discussion following the lab defines enthalpy and entropy. Students then go back into the laboratory where reaction stations are set up around the room. Students rotate through the stations, observe the reaction or process and determining whether it is spontaneous and how enthalpy and entropy are changing. Lecture/discussion compares similar processes (melting wax on a hotplate and melting ice in a room temperature beaker) to emphasize the role of temperature in determining spontaneity, and Gibb’s free energy is introduced.

Our next question is “How hot will a solution of HCl mixed with NaOH get?” After definition of specific heat and a lecture demonstration that iron and water have different specific heats, students are asked to plan a calorimetry experiment:

Experimental overview:

You and your lab partner will measure the enthalpy change per mole of reactants in a reaction of your choice. You may choose either an exothermic or endothermic process, but you should choose a reaction that you know is spontaneous. Decide how you will measure the exact amount of heat gained or lost by the system and write out a detailed set of procedures including chemical hazards and disposal information for all reagents. (Hint: the reading assignment in your text is full of good information.) When your procedures are complete, have them checked by the instructor before proceeding.

Lecture/discussion after the lab emphasizes that enthalpies of reaction need only be

determined by calorimetry once and then are recorded in tables. Students go back into the lab to compare the enthalpy of reacting HCl with solid NaOH vs. the total enthalpy of dissolving NaOH and then reacting it with HCl. Class discussion afterwards applies Hess's law to other examples and introduces the use of enthalpies of formation to determine enthalpy of a reaction. A discussion of the enthalpy changes associated with changes of state follows.

Sections XVII through XX, A First Model of the Atom, An Atomic Model, Planetary Model of the Atom, and Quantum Mechanical Model of the Atom, constitute a unit on atomic structure. The goal of this material is to develop a model of the atom that is useful and consistent with experimental evidence, and is also within the scope of students' mathematical preparation.

In order to encourage students to think about how atomic-scale phenomena can be measured with macro-scale tools, we first ask how big an atom is. In response, students design an experiment to measure the size of an atom. Small groups first generate ideas that they share with the entire class. After a discussion of the relative merits of the various approaches, each lab pair proceeds with the following instructions:

Experimental Overview:

You and your lab partner should spend the next 10 minutes deciding which general strategy you want to use and then write specific procedures to conduct your experiment. Include hazard and disposal information for any chemicals you plan to use, and plan several trials so that you can get a feel for the

precision of your method. When you have recorded your procedures in your lab notebook, and had them approved by your instructor, you may proceed with your experiment.

Lecture/discussion touches on the concept of scientific models in general, and compares atomic volumes measured in lab to several literature values, chosen for their variation rather than their agreement. How can an atom be more than one size? This sets the stage for mystery and counterintuitive results. Experiments of Thompson, Rutherford and Millikan are then presented and information from the periodic table about number of protons, neutrons and electrons is reviewed.

A brief lab experience follows where spectra from different elements are examined.

Students are asked to record the element and their observations in each case.

Lecture/discussion reviews definitions of wavelength and frequency, and presents light as energy quanta. Bohr's model is developed to rationalize the observations that the students made in lab. We then advance to the problems with Bohr's model, define Heisenberg's Uncertainty Principle, introduce Schroedinger's model, describe orbitals and the first four quantum numbers, and develop rules for writing electron configurations.

The eventual goal for Section XXI, Trends in the Periodic Table, is the use of effective nuclear charge to predict and rationalize periodic properties such as atomic radius, electronegativity, ionization energy, etc. Students begin in the laboratory with the following instructions:

Experimental Overview:

For each of the following elements: sodium, potassium, magnesium, calcium, iron, copper, aluminum, phosphorus, sulfur, and iodine, place a small (pea size) chunk in water and observe any results. Use pH paper to test whether the resulting solution is acidic or basic.

Choose another property you want to investigate. Your investigation should include enough elements or compounds that a trend across the periodic table or within a column in the periodic table will be evident. You will have a list of available chemicals, but you must look up hazard and disposal information for any reagents you select. Have your instructor approve your procedures before you begin. Collect your data and observations carefully, and be prepared to present your findings to the class.

Each lab pair presents their results to the class and, of course, many of the properties selected show no real trend in the periodic table, but some do. Why is this, we ask. The instructor uses these examples to begin rationalizing periodic properties, and when the concept of effective nuclear charge seems fairly clear, we progress to periodic properties not measured in lab.

Section XXII, Chemical Bonding, introduces ionic and covalent bonding, writing Lewis structures and the VSEPR model in lecture/discussion. “Given that electrons repel each other, what is the logical shape of methane or ethanol?” is our next question in Section XXIII, Molecular Models. In a laboratory exercise, students use foam balls and toothpicks to determine the logical molecular geometry for a number of compounds. The instructor then reviews the molecular geometry table in the text. Once students are fairly skilled at predicting molecular geometries, the instructor introduces material from Sections XXIV

and XXV, Polar and Nonpolar Molecules and Attraction Between Molecules, which includes uneven sharing of electrons and the use of electronegativity and molecular geometry to determine whether a molecule is polar.

Section XXVI, Properties of Solutions, begins with students measuring several boiling points and looking up several more. Lab pairs are asked to describe patterns in their data and invent theories or rationalizations for those patterns. Students then present their work to the class. Discussion synthesizes students' data patterns and theories, noting inconsistencies. A lecture on intermolecular forces follows. Lecture/discussion then uses comparison of strength of intermolecular forces for the solvent and solute to predict and rationalize solubilities. Colligative properties are then briefly discussed.

Section XXVII, Reactions Which Transfer Electrons, is the last unit because balancing redox equations and stoichiometry with redox reactions are not essential for first semester students and can be abbreviated to a very qualitative discussion if time is short. If abbreviated, the missed material is picked up before we study electrochemistry in the spring. Students begin in the lab with the following instructions:

Experimental Overview:

A number of solutions and solid reagents are available, along with a chart for recording data. With a partner, place the solid reagents in the solutions and observe what happens in each case.

After you have finished recording your observations, discuss your results with another pair. Try to determine what, if any, patterns you see in the data.

The instructor asks “Why do the metals react in some solutions but not others?” The end result, after class discussion, is an activity series. The instructor then presents a strategy for balancing redox equations and works some example stoichiometry problems.

Summary

Approximately half of the syllabi examined indicated implementation of one of three reform measures. Universities and liberal arts colleges were equally likely to indicate innovative activity on the syllabus, although the nature of the activity varied by institutional category. The sample is probably not random and therefore overstates the extent of implementation of reforms. The most common implementation strategy was revising the laboratory curriculum to provide a more investigative structure in introductory courses. This is also the strategy employed by the alternative curriculum presented here.

CHAPTER 5

CONCLUSIONS

Reform recommendations from professional societies indicate disciplinary consensus about the implications of educational research. They are also published in scientific rather than educational journals and are readily accessible to science faculty. Widely endorsed themes for reform were identified in a literature review. This study investigated whether stated directions for reform are actually being translated into changes in undergraduate education. An alternative curriculum was also designed that would implement some of the recommendations identified in the literature review.

Investigation of Reform Implementation in Science Courses

Syllabi were solicited from all 1993 *U.S. News and World Report* first-quartile colleges and universities. A smaller comparison group of third-quartile colleges and universities was examined to verify that colleges and universities with the best reputations and strongest resources are at least as likely as other institutions to implement reform measures.

Telephone interviews of nonrespondents were conducted to ascertain whether the sample was representative of all first-quartile colleges and universities.

As expected, no evidence was found to indicate that curricular reform was more likely at less prestigious institutions. Since the sample of first-quartile institutions was larger than the sample of third-quartile institutions, conclusions are based on data from first-quartile colleges and universities. Most curricular change observed from the syllabi involved enhancing quality in undergraduate laboratories, which was almost always accomplished through use of an in-house laboratory manual. At a ninety-five percent confidence level, there is no significant difference in the percentage of colleges and universities which demonstrate reform implementation on syllabi. However, first-quartile colleges are more likely than first-quartile universities to have a series of investigative labs, rather than just one or two investigative labs.

Telephone interviews of nonrespondents found very little evidence of reform implementation. This indicates a self-selection bias in the sample, despite the fact that no indication was given in the original request for syllabi that innovative programs were being sought.

The two most important results of this survey are that

- 1) there is no significant difference between the proportion of liberal arts colleges and the proportion of universities that have implemented curricular reforms, and
- 2) actual implementation is probably less widespread than the sample would suggest.

We can conclude from the data that less than half of colleges and universities have responded to professional societies' recommendations for science education reform. This is consistent with a recent survey of general chemistry laboratory courses conducted by Michael Abraham of University of Oklahoma.⁶⁵ This survey indicates that seventy-four percent of institutions schedule general chemistry lab for three hours per week, in spite of the high-consensus recommendations for increased laboratory time. Sixty percent of institutions used in-house laboratory manuals, but ninety-one percent of the respondents indicated that students usually followed step-by-step instructions from a lab manual and seventy-nine percent indicated that students do not identify the problems to be investigated. In spite of a large number of recent publications describing innovative laboratories, the results indicate that such innovation in chemistry is confined to a few institutions. This study does not tell us whether chemistry education is similar to other sciences.

While reform is not as prevalent as we might hope, the analysis of syllabi does indicate that colleges are more likely to display extensive reform than universities and that reform is most likely to be visible in the laboratory portion of introductory courses. Apparently, colleges and universities are responding to the call for change in science education by transforming the laboratory first. One can presume that this is at least partially a result of practical considerations. Space and personnel limitations are probably more significant barriers to increasing the amount of time spent in laboratory (strategy 1 in this study) than

they are to altering the content of those laboratory experiences (strategy 2). Similarly, large lecture sections do not readily facilitate evaluation or development of students' communication skills (strategy 3). It would be interesting to investigate whether discipline-specific communication skills are being taught in the small group laboratory setting.

Although changes in the content of laboratory experiences may not require a radical change in the university structure, there are still difficult administrative issues to resolve. Teaching assistants for investigative laboratories will require more sophisticated training than graduate students have traditionally received. Introducing a project-oriented format increases reagent and materials preparation time, and may represent a significant burden for stockroom and laboratory managers, as well as increasing costs. Since not every detail is planned in advance, safety training and vigilant safety rule enforcement is critical. Despite major practical issues like these, some colleges and universities are changing the way they do undergraduate and even introductory laboratory. These results are heartening and indicate serious attention to effectiveness of undergraduate education.

College Chemistry Curriculum

A first-semester curriculum for general chemistry is presented in Appendix A. The arrangement and number of topics is quite traditional, but an investigative laboratory is

central to the course. The results from two semesters of implementation have been very positive. As an example, measuring the ideal gas constant frequently is assigned in “cookbook-style” lab manuals, and it never works very well for freshmen. This leads to two negative outcomes: students are frustrated that they cannot get the right answer, and students think that the goal of an experiment is to get a predetermined right answer.

This curriculum asks students to measure the ideal gas constant, but they design their own experiment to do so. Students still do not measure R very accurately, but instead of being frustrated, they are encouraged to assess their experimental design and suggest how it might be improved. They still have a concrete opportunity to measure the important variables in gas behavior and get a handle on what R means, but without the negative outcomes previously mentioned.

Another positive outgrowth of an investigative approach to lab is that students learn from each other as they see someone across the room doing a different experiment. It is possible that they actually learn more descriptive chemistry than in a traditional laboratory course, although this assertion has not been tested.

It is also interesting that academically successful students are not necessarily the only strong performers in this type of laboratory setting. It seems that academically stronger students learn to appreciate creative contributions from weaker students, which is an unintended but positive outcome. Again, no standard values and attitudes measure has

been used to verify these qualitative observations.

The primary shortcoming of this curriculum is the intensive staff time required. The maximum lab section size with which this curriculum has been used is eighteen students, who usually work in pairs. Mentoring more than nine projects at the same time would be very difficult for one person.

The results in terms of student learning and self-confidence in the laboratory are quite satisfactory. Comparing student achievement in this integrated, investigative laboratory with that of a traditional lecture plus a more investigative laboratory, such as Abraham's *Inquiries Into Chemistry* is an important next step.

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APPENDIX A

FIRST-SEMESTER GENERAL CHEMISTRY COURSE OUTLINE

Elements, Compounds and Mixtures

Objectives: Upon completing this material, the successful student should be able to

- 1) distinguish between an element, a compound and a mixture,
- 2) begin formulating a strategy for eliciting patterns from data, and
- 3) relate the definitions for atom and molecule to those for element, compound and mixture.

Reading assignment: Umland, Chapter 1, pp 2 - 11

Chemistry tends to focus on very tiny pieces of the world: atoms and molecules.

Atoms are very small indeed: a 150 pound person contains 2.7×10^{28} atoms.¹ We will start our examination of the world at this microscopic level by trying to answer three questions:

- 1) What is an atom?
- 2) How much do atoms weigh?
- 3) How can we measure the number of atoms if they are too small to count?

Chemical Inquiry - What is an atom?

Each pair of students will have several jars that contain an "element", a "compound" and a "mixture." (HINT: Labels for elements will only have one capital letter.)

Examine these substances and write down as many observations as possible. As a class, we will develop definitions for these three terms based on our observations.

¹Assuming an average person is 62 % by weight hydrogen, 26% carbon, 10% oxygen, 1.5% nitrogen and 0.5% other elements, mainly calcium and phosphorus. Martini, Frederic. *Fundamentals of Anatomy and Physiology*, 2e. Prentice-Hall, NJ, 1992. p. 29

Your definitions:

Elements -

Compounds -

Mixtures -

Inquiry Discussion

Elements:

Compounds:

Mixtures:

Atoms are individual units. Molecules are collections of atoms chemically linked together. It is important to remember that molecules have very different properties than their individual atoms had separately. Properties of atoms do not change when they are just mixed together without becoming chemically linked.

Practice

Classify the following substances using the definitions we have developed.

cola	blood	Kool-Aid
water	air	liquid nitrogen
dry ice	24 carat gold	14 carat gold
vitamin C	plutonium	mercury

Give three examples not discussed so far for each of the following classifications.

Element

Compound

Homogeneous mixture

Heterogeneous mixture

Application

A dark blue substance is heated in a test tube. A vapor rises from the substance as it is heated. After several minutes of heating the substance turns light blue. Is the dark blue compound

- a) an element b) a compound
- c) a mixture d) not enough information to determine

Is the light blue substance

- a) an element b) a compound
- c) a mixture d) not enough information to determine

Naming Compounds

Objective: Upon completing this material, the successful student should be able to

- 1) define a number of terms used in our textbook, and
- 2) name inorganic compounds.

Reading Assignment: Umland, Chapter 1, pp 4 - 24

Ticket to Class

Bring definitions for the following to class:

physical change	chemical change
metal	nonmetal
periodic table	atomic number
molecular compound	ionic compound

Inorganic Nomenclature

The appendix of your text provides rules for nomenclature. Refer to them when you are unsure of the name of a compound.

Common Polyatomic Ions

These polyatomic ions and their charges should be memorized. A complete list of polyatomic ions, including less common ions, is found in Appendix B of your text.

$\text{C}_2\text{H}_3\text{O}_2^-$	acetate	CO_3^{2-}	carbonate
CN^-	cyanide	PO_4^{3-}	phosphate
SO_4^{2-}	sulfate	SO_3^{2-}	sulfite
OH^-	hydroxide	ClO_4^-	perchlorate
NO_3^-	nitrate	NO_2^-	nitrite
S^{2-}	sulfide	NH_4^+	ammonium

Initial Strategy for Writing Formulas

All compounds must have a net charge of zero.

If the metal name is followed by a Roman numeral, that defines the charge of that metal ion. Adjust the number of negative ions so that the total charge is zero.

If a compound contains a polyatomic ion, the negative charge is defined. Adjust the number of positive ions so the charge is zero.

Application

Twenty different compounds are arranged throughout the room. Ten are identified by chemical symbols; ten are identified by full name.

Write the names of the compounds that are identified by symbol. Make sure you have ten DIFFERENT compounds. (There may be more than one container of a compound.)

- 1.
- 2.
- 3.
- 4.
- 5.
- 6.
- 7.
- 8.
- 9.
- 10.

Write formulas for the compounds that are identified by name. Make sure you have ten DIFFERENT compounds. (There may be more than one container of a compound.)

- 1.
- 2.
- 3.
- 4.
- 5.
- 6.
- 7.
- 8.
- 9.
- 10.

You should be able to complete 1- 11, 13 and 14 of “Stop and Test Yourself” in Chapter 1.

Safety and Disposal

Objective: Upon completing this material, the successful student should include chemical hazard and disposal information in any set of laboratory procedures.

Chemical hazard data tells you whether and which special precautions you should take in working with a particular substance. We will always wear safety glasses, and will take other precautions as necessary. Hazard and disposal information can be found in a variety of sources, several of which will be on reserve in the library.

NFPA labeling is used to indicate to chemists (and fire fighters) the risks of a particular chemical. The blue diamond is health related risks, red indicates flammability, yellow describes reactivity and white is for special precautions.

This is one definition of the NFPA labeling system.² Variations on this definition can be found posted in the chemistry hallway and the laboratory. These various definitions have many similarities, but are not identical. In summary, a health rating of two does not have a precise meaning. Your best bet is to look up each chemical.

Blue

- 4 Death or serious injury after short exposure
- 3 Serious injury possible after short exposure
- 2 Intense exposure may cause injury (usually temporary if medically treated)
- 1 Only causes irritation even without medical treatment

Red

- 4 Burns readily and is easily spread through air
- 3 Can be ignited at room temperature
- 2 Must be heated before it will ignite
- 1 Must be preheated before it will burn
- 0 Will not burn

Yellow

- 4 Explosive decomposition at normal temperatures
- 3 Explosive decomposition when heated
- 2 Unstable and reactive but does not spontaneously explode
- 1 Unstable and reactive when heated
- 0 Stable even under fire conditions; does not react with water

²Weiss, G. *Hazardous Chemicals Data Book*, 2e. Noyes Data Corp., Park Ridge, NJ, 1986. pp. 11-12

Disposal

Nontoxic aqueous solutions can be rinsed down the drain, although very acidic or very basic solutions should be neutralized before disposal.

Organic solvents (except for ethanol) should be saved in an organic waste container. We will try to minimize our use of organic solvents.

Moderately toxic metals should be precipitated out of solution before the aqueous solution is disposed of in the sink. The solid precipitate can be **LABELED** and saved for another experiment. If it is contaminated or otherwise unsuitable for reuse, it can be sealed in a jar and thrown away in the trash.

We will generally not use highly toxic metals, such as lead and mercury, in lab experiments.

Assignment

Five chemicals from our stockroom and five household chemicals are available in the lab.

For each, determine

- 1) what potential hazards they present,
- 2) what protective measures are required to use them safely, and
- 3) how they should be disposed of.

Atomic Masses

Objective: Upon completing this material, the student should

- 1) use appropriate significant digits in laboratory work,
- 2) include instrument calibration checks in any experimental procedure, and
- 3) be able to explain the experimental procedures s/he has designed.

Reading assignment: Umland, Chapter 1, pp 15- 19; Chapter 2, pp 45-51 and 59-64.

Ticket to Class

Bring a summary of the rules for determining significant digits and your signed lab safety contract.

Before we actually perform the next chemical inquiry, we need to define weight and mass, significant digits, and precision and accuracy.

Weight and Mass

Weight is the force of gravity on an object. Mass is how much matter an object has. Astronauts weigh less on the moon and weigh nothing in outer space. Their mass does not change.

Significant Figures

Significant digits or significant figures tell you what numbers have been recorded from a measurement. The number of significant digits often tells you something about how the measurement was made. You would record a different number of significant figures if you weighed an object on a bathroom scale rather than a laboratory scale.

Rules for Significant Digits

Precision vs. Accuracy

Precision is a measure of how consistent measurements are.

Accuracy is a measure of how close they are to the true value.

Most of the time we don't know the true value and have to guess at how accurate our measurements are.

Chemical Inquiry - How much do atoms weigh?

During the next class period, you will convert magnesium (Mg) to magnesium oxide (MgO). Let's say you started with a pound of magnesium (Mg). When you convert it to magnesium oxide, you are linking an oxygen atom to every magnesium atom. This added oxygen should weigh something, so your end product will weigh more than a pound. How much do you think it would weigh?

Experimental overview:

Many elements will combine with oxygen from the air when heated. Magnesium, a solid metal, will combine with oxygen to produce magnesium oxide (MgO), but it also combines with nitrogen to form magnesium nitride (Mg₃N₂). Water will convert magnesium nitride to magnesium hydroxide, releasing ammonia (NH₃) into the air. Heating magnesium hydroxide will form magnesium oxide, a white or gray solid. We will use these reactions to compare the mass of a magnesium atom with the mass of an oxygen atom, since MgO has equal numbers of each.

Together with your lab partner, design a method to fully convert magnesium to magnesium oxide. You will need to determine the final weight of the oxygen used in the compound, and the ratio of Mg mass to O mass. Make your procedures as detailed as possible and be sure to include hazard and disposal information for chemicals. Triple space your procedures so that you can add information.

Assignment

Bring completed procedures in your notebook to the next class, with space for corrections.

Practice

Umland, chapter 2, pp. 49 - 51 #19, 20

Application

Umland, Chapter 2, p. 72 #97

Measurements and Data

Objective: Upon completing this material, the successful student should be able to

- 1) record data from a variety of instruments to the correct number of significant digits,
- 2) use the median and average of a data set in analyzing data,
- 3) articulate an analysis of data collected,
- 4) evaluate experimental procedures in light of experience, and
- 5) use the periodic table and a list of the elements to determine atomic masses.

Reading Assignment: Umland, Chapter 2, pp 45 - 51 and 59 - 66

Recording Measurements

Keep the following in mind:

- 1) The last digit you record contains uncertainty.
- 2) The last digit of an electronic readout is always uncertain (the instrument estimates for you).
- 3) If the instrument is not functioning well, a digit well before the last digit of an electronic readout may be uncertain. This is a judgement call, but you must have some rationale to defend your judgement.
- 4) When reading a scale (not a digital display), estimate the last digit.

Chemical Inquiry:

If you have had your notebook procedures approved, proceed with the experiment you planned. BE SURE TO TURN IN THE MASS OF OXYGEN THAT YOU DETERMINE BEFORE YOU LEAVE.

Inquiry Discussion

Class median:

Class average:

Atomic weights listed on the periodic table were determined using the same principles you used in comparing the mass of magnesium and oxygen. Today all atoms are compared to the mass of ^{12}C . Mass of atoms is measured in atomic mass units. By definition, ^{12}C has a mass of 12.0000 amu, so one amu is exactly 1/12 of the mass of ^{12}C .

Atomic number defines the number of protons. Atomic mass is the number of protons plus the number of neutrons.

Atoms with the same number of protons and different numbers of neutrons are isotopes of an element.

The number of electrons equals the number of protons in a neutral atom.

Practice

Look up the atomic symbol and atomic mass for each of the following:

sodium	phosphorus
nitrogen	mercury
chlorine	calcium
iron	iodine
lead	radon
sulfur	silicon

Umland, Chapter 2, pp. 62 - 66 # 31 - 38

Application

Umland, Chapter 2, p. 73 #113

Assignment

1) What parts of your experimental procedures were especially efficient or accurate? What parts of your experimental procedures would you change if you were planning to perform this inquiry again? Turn in a one to two page evaluation of your experiment by _____.

2) When you have completed your analysis of data, turn in your notebook.

Metric Units

Objective: Upon completing this material, the successful student should be able to

- 1) create and use conversion factors, and
- 2) define the most common metric units for mass, volume, density, temperature and time.

Reading Assignment: Umland, Chapter 2, pp 38 - 44 and 52 - 58.

Conversion Factors

Conversion factors express the same quantity in different units. For example, 12 inches is the same length as 1 foot. 16 ounces of cheese is the same as 1 pound of cheese.

$$12 \text{ inches} = 1 \text{ foot}$$

$$1 \text{ pound} = 16 \text{ ounces}$$

$$\frac{12 \text{ inches}}{1 \text{ foot}} = 1$$

$$\frac{16 \text{ ounces}}{1 \text{ pound}} = 1$$

To find the number of ounces from the number of pounds, I multiply by the fraction " $\frac{16 \text{ ounces}}{1 \text{ pound}}$ ". This amounts to multiplying by 1, which does not change the quantity, just the units.

EXAMPLE 1

$$0.75 \text{ pounds} \left(\frac{16 \text{ ounces}}{1 \text{ pound}} \right) = 12 \text{ ounces}$$

12 ounces and 0.75 pounds are the

SAME

EXAMPLE 2

Jessica and Rolff are planning a trip to Colorado for spring break and are trying to figure out how much money to take with them. Rolff has offered to drive his family's Volvo. Jessica estimates from a road atlas that they are 560 miles from the slopes. Rolff says his car gets 20 km per liter. To plan how much they will spend on gas, they need to know how many gallons of gas they will use.

$$\frac{560 \text{ miles}}{20 \text{ km/L}} \neq \text{number of gallons of gas needed}$$

One mile is the same as 1.6 kilometers. One gallon is the same as 3.785 liters. When

written as fractions, these are called conversion factors.

$$1 \text{ mile} = 1.6 \text{ km}$$

$$1 \text{ gal} = 3.785 \text{ L}$$

$$\frac{1 \text{ mile}}{1.6 \text{ km}} = 1 \quad \frac{1.6 \text{ km}}{1 \text{ mile}} = 1$$

$$\frac{1 \text{ gal}}{3.785 \text{ L}} = 1 \quad \frac{3.785 \text{ L}}{1 \text{ gal}} = 1$$

Multiplying by a conversion factor is the same as multiplying by one. So when we multiply by a conversion factor, we are not changing the value of any measurements, but are only changing the units.

IMPORTANT: When using conversion factors, treat units just like numbers or variables.

- 1) Only cancel units that appear in the numerator and the denominator.
- 2) Only cancel units that are exactly alike.

If the units of your answer do not seem to make sense, that is a sure sign that you have a problem in your calculation.

Metric Units

Most scientific measurements use metric units.

<u>Quantity</u>	<u>SI Unit</u>	<u>Common Unit in Chemistry</u>
Volume	cubic meter	liter (L)
Mass	kilogram	gram (g)
Density		grams/mL or grams/cc
Temperature	Kelvin	degree Celsius (°C)
Time	second	second (s)

SI prefixes for metric units are listed on the inside back cover of your text. You will be held responsible for knowing nano, micro, milli, centi, deci, kilo and mega.

Practice

Umland, Chapter 2, pp 42 - 45 #3 - 11; pp 54 - 59 # 23 - 30; pp 73 - 74 #109 and 121

You should be able to complete # 1 - 12 of “Stop and Test Yourself” in Chapter 2.

Chemical Moles

Objective: Upon completion of this material, the successful student should be able to

- 1) define the term "mole," and
- 2) use the periodic table to determine the mass of a mole of a substance.

Reading assignment: Umland, chapter 2 pp 65-67; Chapter 3 pp 79-81

Counting atoms

By looking at the periodic table, we can see that 1 gram of hydrogen, 12 grams of carbon and 32 grams of sulfur all have the same number of atoms. This number of atoms is called a mole. How many atoms do you think are in a mole? What kind of experiment could you do to find out?

In groups of four, outline an experiment to determine the number of molecules in a mole. You may pretend that any materials are available to you, but bonus points will be awarded to groups that use materials available in our labs.

Summary: A mole is just a name for a large number (6.022×10^{23}), just like a dozen is a name for 12 of something.

The periodic table gives atomic masses. Appropriate units are either amu (the average mass of one atom), or grams/mole (which means the number of grams that 6.02×10^{23} atoms weigh).

So how much does one magnesium oxide weigh? From the periodic table,

$$\begin{aligned}\text{mass of MgO} &= \text{mass of Mg} + \text{mass of O} \\ &= 24.31 \text{ amu} + 16.00 \text{ amu}\end{aligned}$$

$$\text{MgO} = 40.31 \text{ amu} \quad \text{or} \quad 40.31 \text{ g/mol}$$

Practice:

Umland, Chapter 2, pp 71- 74 #95 and 99; Chapter 3, pp 80-81 #1-3

Application:

There are five substances in lab. Write the formula and calculate the molecular weight for each.

You should be able to complete “Stop and Test Yourself” from chapter 2.

³Chemists have known since the 1800's that 2 grams of H₂ and 16 grams of O₂ had the same number of atoms. The mole was first measured by Loschmidt in 1865. (Virgo, S.E., "Loschmidt's Number", *Science Progress*, 1938, 27, pp 634-49)

Occasionally new technology provides a more accurate measure and the accepted value of the size of the mole changes slightly. Redetermination of the atomic mass of silicon in 1992 led to a slight correction in the accepted value for the mole. The first four digits, which is all we will ever need to use, however, are well established. (J.R. DeLaeter, P. De Bièvre, and H.S. Peiser, "Isotope mass spectrometry in metrology," *Mass Spectrometry Review*, 1992, 11(3), 193-245)

Chemical Formulas

Objective: Upon completion of this material, the successful student should be able to

- 1) discriminate between molecules that are similar, but not alike,
- 2) differentiate between an empirical and molecular formula, and
- 3) use percent composition data to find an empirical formula or *vice versa*.

Reading Assignment: Umland, Chapter 3 pp 94 -103

When we look at the periodic table, there are surprisingly few elements. Only about a third of the known elements make up everything found in this room, including the people⁴. There seems to be an endless array of colors, textures, and odors all around us. How can so few elements produce such wonderful variation? Let's look at how elements form compounds.

The same elements can be put together in different ways to form different chemical compounds. Examples include:

O₂ (molecular oxygen) and O₃ (ozone)
C₁₀H₈O₄ (polyester) and C₆H₁₂O₆ (glucose)
C₃H₇O₂N (L-alanine) and C₃H₇O₂N (D-alanine)

The properties of any compound depend on which elements are present and how the atoms are arranged. Deciding how atoms are arranged is fairly complex, but we already have the tools to determine how much of a certain element is in a compound.

Chemical Inquiry - Determination of a Chemical Formula

Experimental Overview:

We used the weights of a reactant (magnesium) and a product (magnesium oxide) to determine the weight of the other reactant (oxygen). To accomplish this we had to know the formula of magnesium oxide (MgO, one magnesium for every oxygen). This time you will use known weights of the reactants to determine the formula of the compound.

⁴This assumes that all of the noble gases are found in the air, that all of the nonmetals plus silicon are likely to be present, and that sodium, magnesium, potassium, calcium, aluminum, titanium, chromium, manganese, iron, copper, cobalt, nickel, zinc, zirconium, cadmium, molybdenum, silver, tin, gold, mercury, and lead are found in people or everyday objects.

You will prepare a new compound by mixing together two solutions. Then you will determine how much of each substance is in the new compound.

A solution of iron will be provided, labeled with the number of grams iron per mL of solution. Sodium hydroxide (a solid) is soluble in water. When iron and hydroxide solutions are mixed, a precipitate (solid) will form.

With your lab partner, develop a detailed set of procedures to determine the mass of each substance in your final compound. This will allow you to calculate the formula. Make sure you consider how you will know whether all of your reactants were used to form the solid. Include chemical hazard and disposal information for any reagents you plan to use. Have the instructor check your completed procedures.

Express the amount of each substance in the compound as a percent of the total mass. Determine a formula from the masses you measured. Turn in your data.

Inquiry Discussion:

Percent composition can be used to determine a chemical formula. The formula we get is called an empirical formula because it only tells us what the ratio of atoms is to each other; it does not say anything about the way that they are arranged.

Sometimes two very different compounds can have the same empirical formula. For example, acetic acid ($\text{C}_2\text{H}_4\text{O}_2$) and glucose ($\text{C}_6\text{H}_{12}\text{O}_6$) have the same empirical formula (CH_2O).

If we have the total molecular weight of the compound and the empirical formula, we can determine the exact number of each kind of atom in the molecule.

Practice:

Umland, Chapter 3, pp 94-103 #11 - 17 and 21; pp 105 #59 and 61

Application:

pp 105 -108 # 67, 77, 81, and 93

You should be able to complete #2, 3 and 10 - 15 of “Stop and Test Yourself” in Chapter 3.

Chemical Reactions

Objective: Upon completion of this material the successful student should be able to

- 1) distinguish chemical reactions from physical changes,
- 2) explain the significance of the first law of thermodynamics to chemistry,
- 3) identify the category in which a particular reaction belongs and begin to predict products, and
- 4) balance chemical equations.

Reading Assignment: Umland, Chapter 3, pp 81-84

The subject of chemistry is essentially organized around studying the properties of different substances and how substances interact. Besides satisfying our curiosity about why some things catch fire easily or why oil and water don't mix, this way of thinking gives us a model for understanding elaborate systems which exhibit change. Living organisms, for example, can be thought of as a set of complex, organized, interdependent chemical reactions.

What are some examples of chemical reactions occurring in this room? Outside?

A chemical reaction is

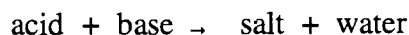
There are as many different reactions as there are substances around us, but they all follow a few basic rules. We will only discuss the first of these rules for now.

Law of Conservation of Mass and Energy : $E=mc^2$ tells us that a tiny bit of mass is the same as a huge amount of energy. Since we will always deal with relatively small amounts of energy, we will never observe a change in mass during a reaction. The total masses before the reaction will always equal the total masses after the reaction.

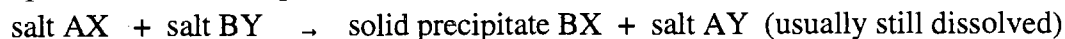
Chemical equations are a shorthand way to write out what happens during a reaction. Things we start with are called reactants. Things that are made are called products.

Common Reaction Types

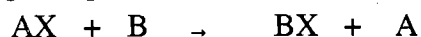
Acid - Base



Precipitation (Double Displacement)



Single displacement



Combustion



Decomposition



Chemical Inquiry - Writing and balancing equations

Labeled reagents have been prepared for you. With your partner, mix these reactants together and observe the results. Try to identify the reaction, write an equation for the reaction, and balance your equation. For each double displacement reaction, try to identify the precipitate. When you are done, check your results with another group.

Inquiry Discussion

Chemical equations describe what atoms are doing. Atoms do not disappear into thin air or pop up out of nowhere, so we must always have exactly the same number of each kind of atom on both sides of the equation.

In balancing equations, we can only change the coefficients (numbers in front of each compound). We can NEVER change subscripts within a compound, because that would turn it into a different chemical.

Practice

Umland, p 84 # 4

Application:

Umland, p 107 # 73

Calculations in Chemistry

Objective: Upon completion of this material, the successful student should be able to

- 1) translate a written description of a reaction to a chemical equation or *vice versa*,
- 2) utilize the concept of a limiting reactant to predict the amount of products in a particular chemical reaction and whether any reactants will be left over, and
- 3) explain why the number of moles, not the mass of the reactants, determines which is the limiting reactant.

Reading Assignment: Umland, pp 84 - 93

Ticket to Class

Look at your data for the inquiry where you made iron (III) hydroxide. Write down in the space below how much of each reactant you started with and whether each reactant was completely used up.

Reactant formula 1) _____ 2) _____

Mass of reactant 1) _____ 2) _____

Reactant used up? _____

Probably, one of your reactants was completely used to form the precipitate and some of the other reactant was left over. The reactant that was completely used up is called the limiting reactant.

This is a concept you use every day. Let's say you have a shop that rebuilds bicycles. You need one frame for every two wheels. If you have 10 frames and 8 wheels, how many bicycles can you make?

We can write a "chemical reaction" for your business.

1 frame + 2 tires → 1 bicycle

This reads either "one frame and two tires yields one bicycle" or "one mole of frames and two moles of tires yields one mole of bicycles."

We can use this same reasoning to predict how much product we get out of a chemical reaction and which reactant will be left in solution.

Write a balanced equation for the reaction you performed in the space below. When you are done, check your answer with someone else.

Write a sentence that contains all the information that your balanced equation contains. When you are finished, compare your sentence with a neighbor's.

The equation for precipitation of iron (III) hydroxide says that one iron (III) nitrate reacts with three sodium hydroxides to produce one iron (III) hydroxide and three sodium nitrates.

It also says that one mole of iron (III) nitrate reacts with three moles of sodium hydroxide to produce one mole of iron (III) hydroxide and three moles of sodium nitrate.

A solution containing five moles of iron (III) nitrate and another solution containing 13 moles of sodium hydroxide are mixed together. Iron (III) hydroxide precipitates from the solution and is removed by filtration. The liquid from filtration is tested by adding a few more drops of iron (III) nitrate. Will a precipitate form?

The equation reads: one iron (III) nitrate reacts with three sodium hydroxides to produce one iron (III) hydroxide and three sodium nitrates. It does NOT say that 1 gram of one thing and 3 grams of another give 1 gram of products.

Any time you are using chemical equations, convert all masses to moles.

50 grams of iron (III) nitrate and 35 grams of sodium hydroxide are mixed together in solution. Iron (III) hydroxide precipitates from the solution and is removed by filtration. The liquid from filtration is tested by adding a few more drops of iron (III) nitrate. Will a precipitate form?

Assignment:

Bring the following completed problems to class:

- 1) A solution of 75 grams of calcium chloride (CaCl_2) is mixed with a solution containing 100 grams of sodium phosphate. When the precipitate is removed, will there be any calcium ions left in the solution?
- 2) Calculate the expected yield of product from the data you recorded before class.

Practice:

Umland, p. 105 # 65

Using Stoichiometry

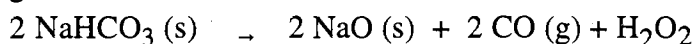
Objectives: Upon completion of this material, the successful student should be able to

- 1) use balanced equations and observation in chemical analyses, and
- 2) solve common stoichiometric problems.

Reading Assignment: Review Umland, pp 86 - 94

Ticket to Class

Baking soda is sometimes suggested as a kitchen fire extinguisher. Is it possible to create dangerous byproducts by overheating baking soda? In the next chemical inquiry, you will be given a sample of sodium bicarbonate (baking soda), which decomposes when it is heated. Write at least two balanced chemical equations for possible decomposition products of NaHCO_3 . Both your textbook and the CRC can be helpful in determining whether a product you propose is a known compound. An example is given below:



Chemical Inquiry - Stoichiometric relationships

Experimental overview:

A list of possible reactions for the decomposition of NaHCO_3 is available. You should plan a strategy to help you determine which, if any, of these equations best describe decomposition of NaHCO_3 .

When you have a strategy recorded, discuss it with a neighboring lab group and then write out detailed procedures. You may perform a qualitative investigation before you plan detailed procedures. Since you do not know what your products are, save the products for disposal after they have been identified.

When you have finished your investigation, record which of the possible equations best describe decomposition of NaHCO_3 , and your reasoning for your choice. Be prepared to turn this in next class period.

Inquiry Discussion

When we use reactions in the laboratory, we usually want to know how much of one reagent to add in order to completely react with another. For one thing, it would be wasteful to add a lot more of one reagent than the reaction can use and then have to throw it away. In fact, it is often more expensive to dispose of a chemical than it is to purchase it!

Some common types of stoichiometry problems:



- 1) How much of B do I need to completely react with A? How much of C do I expect to get?

Example: Umland #102 p. 109

- 2) I need 50 grams of C. How much A and B should I use?

Example: Umland # 75 p. 107

3) I have 35 grams of A and 75 grams of B. How much C will that produce? Which reagent will be left over and how much?

Example:Umland # 83

4) I calculated that this reaction would produce 50 grams. When the actual reaction was performed, only 25 grams of C were recovered. What was the percent yield?

Example:Umland # 83

Strategy for stoichiometry problems:

- 1) start with a balanced equation
- 2) convert all masses to moles
- 3) determine which reactant is limiting
- 4) use unit analysis to set up a ratio between moles of one reagent and moles of another
- 5) convert moles back to grams, if necessary

Practice

Umland, Chapter 3, pp. 86 - 93 # 5 - 10

Application

Umland, Chapter 3, p. 105 # 63; p. 107 # 71 and 85

You should be able to complete “Stop and Test Yourself” in chapter 3.

Reactions in Solution

Objectives: After completion of this material, the successful student should be able to

- 1) define and use various units of concentration, and
- 2) extend stoichiometric skills to reactions in solution.

Reading Assignment: Umland, chapter 4, pp 112 - 116 and 130 - 145

Ticket to Class

Define concentration in the space below, specifying how concentrated and dilute solutions are different. Be prepared to explain your definition.

The most common unit of concentration is molarity (M). Molarity is defined as the number of moles of solute divided by liters of solution.

Practical questions encountered in the laboratory often relate to preparing and using solutions. The following examples indicate some of the most common lab situations.

- 1) How do I make 500 mL of a 0.25 M solution of NaCl?
- 2) I have a large stock bottle of concentrated (12 M) HCl and I need 2 L of 2 M HCl. How do I make it?

3) The directions say to add 0.3 moles of NaOH. How many mL of 2 M NaOH should I use?

4) I am mixing 150 mL of 2 M KCl and 250 mL of 1 M NaF. What are the final concentrations?

We will need the same information when we work with reactions in solution as we do when we use solid reagents. The calculations for solutions are just like those with solid reagents, except your starting point is a volume and a concentration, instead of grams.

Strategy for solution stoichiometry problems:

- 1) start with a balanced equation
- 2) convert all volumes and concentrations to moles
- 3) determine which reactant is limiting
- 4) use unit analysis to set up a ratio between moles of one reagent and moles of another
- 5) convert moles back to volume and concentration, if necessary

Practice

Umland, pp. 132 - 142 # 23 - 32

Application

Umland, p. 147 # 71, p. 148 # 75, p. 149 # 85, 89, and 93

You should be able to complete #1 - 8 and 15 of "Stop and Test Yourself" in chapter 4.

Properties of Gases

Objectives: After completing this material, the successful student should be able to

- 1) explain the relationships described by the ideal gas law and the conditions under which that law is valid,
- 2) define standard temperature and pressure,
- 3) use partial pressure to determine the amount of a particular gas in a mixture of gases, and
- 4) extend stoichiometric skills to reactions involving gases.

Reading Assignment: Umland, chapter 5, pp 154 - 179

When we pump air into tires or out of vacuum tubes, we are moving gas molecules from one place to another. The difference between individual molecules of gases and individual molecules of solids or liquids is simply the distance from their neighbors and how fast they move.

Gas molecules are very far apart and are moving very quickly, whereas solid and liquid molecules are closer together and tend to move more sluggishly.

The things that are easiest to measure for gases are pressure and temperature. Pressure is usually measured in mm of mercury, and often converted to atmospheres (atm). Temperature is measured in Celsius and converted to Kelvin.

What do you already know about gas behavior? In groups of three or four, answer the following questions:

1) What happens to a gas when you heat it up? (Think about hot air balloons, and try to come up with another example that is consistent with your answer.)

2) What happens if you heat a gas in a closed container? (Consider the difference in your tires on a cold winter morning, and at the end of a long drive. Give another example that is consistent with your answer.)

3) What happens if you change the amount of gas in a closed container? (How do a brand new aerosol bottle and an almost empty one behave differently? Give another example that is consistent with your answer.)

Ideal Gas Law

Chemical Inquiry - Ideal Gas Constant

We will determine the value of the ideal gas constant, R . In groups of four, devise an experiment that will allow you to measure the quantities necessary to determine the ideal gas constant. When you have outlined a set of procedures, consult the troubleshooting checklist to correct any major omissions, and then have the instructor check your final procedures before you begin work.

Before calculating R , convert your volume to liters and your temperature to Kelvin (absolute temperature; $K = 273 + ^\circ C$). Leave pressure in mm Hg. When you have completed your experiment, turn in your results.

Inquiry Discussion:

$PV = nRT$ is called the ideal gas equation. The units on R determine the units for all the other quantities. In this class we will use $R = 0.0821 \text{ L atm/mol K}$. This means volume is measured in liters, n is the number of moles of gas, pressure must be converted to atmospheres and temperature must be converted to K.

The ideal gas equation gets its name because it is only accurate with certain gases under “ideal” conditions. As long as the temperature is fairly high and pressures are fairly low, many gases behave ideally.

Because gas volume depends on the temperature and pressure, in order to have a common reference point, chemists have defined a standard temperature and pressure (STP). The temperature and pressure that were arbitrarily chosen were 0°C (273 K) and 1 atm .

We can use the ideal gas equation in stoichiometry problems like the following.

1) Nitrous oxide (laughing gas) can be prepared by heating NH_4NO_3 which decomposes to N_2O (g) and H_2O (g).

How would you separate the two gases?

How much ammonium nitrate should you start with if you want to produce 0.5 L of nitrous oxide at 20°C and 700 mm Hg ?

2) Hydrogen gas is collected in a balloon for use in a demonstration by dissolving 50 grams of zinc in 100 mL of 0.1 M HCl . Room temperature is 20°C and the barometer reads 640 mm Hg . What volume of gas will be collected? (Important to know in choosing the size of your balloon!)

Dalton's Law of Partial Pressures

It is impossible to store gases for a long time in a container like a balloon. If we are collecting gases for later use, we will want to use some sort of solid container, like a glass flask or test tube.

One common strategy is to collect gases over water. This system is useful when water vapor mixed in with the gas of interest will not interfere with the use of that gas.

Dalton's Law of partial pressures says that mixtures of ideal gases do not interact, so that the total pressure of a gas mixture is the sum of the pressures of the individual gases. For example, if 1 L of nitrogen at 1 atm is added to 1 L of oxygen at 2 atm in a closed, rigid container, the total pressure will be 3 atm.

Because number of moles and pressure are directly proportional, Dalton's Law also works for moles of gas. For example, 0.5 mole of helium in a volume of 22.4 L has a pressure of 0.5 atm. 1 mole of hydrogen in a 22.4 L container has a pressure of 1 atm. If the two are placed in the same 22.4 L container, the total number of moles is 1.5 and the total pressure is 1.5 atm.

One practical application of Dalton's Law is that it allows us to calculate how many moles of water vapor and how many moles of another gas are in a mixture that has been collected over water.

The amount of water vapor in the air or any ideal gas mixture is determined by the temperature. These amounts have been carefully measured and tabulated.

Practice

Umland, pp 159 - 161 # 3 - 8, pp 165 - 179 # 10 -30

Application

Umland, pp 191 # 87, 91, 93

You should be able to complete #1 - 9 and 15 of "Stop and Test Yourself" in chapter 5.

Kinetic Molecular Theory of Gas Behavior

Objectives: After completing this material, the successful student should be able to describe the kinetic molecular model of gas behavior and explain and use Graham's Law of diffusion.

Reading Assignment: Umland, chapter 5, pp 181 - 186

Ticket to Class

A half filled balloon contains enough air to occupy 500 mL at 0°C. The temperature is increased and the balloon's final volume is 600 mL. Calculate the final temperature and draw a picture of the air inside the balloon before and after heating.

The kinetic-molecular theory of gases is an attempt to explain gas behavior in terms of the behavior of individual atoms. It is especially useful for thinking about temperature and pressure.

Assumptions of Kinetic-Molecular Theory

- 1) Gas molecules are very small in comparison to the space between them.
- 2) Each gas molecule is moving very fast.
- 3) Gas molecules do not attract or repel each other if they collide.

Applying KMT

Pressure is explained by the kinetic-molecular model as the force of the gas molecules striking solid objects (including the walls of a container).

Temperature, in the kinetic-molecular theory, is directly related to the average kinetic energy of the gas. Gases with a high average kinetic energy have high temperatures; gases with low average kinetic energy have low temperatures.

An increase in pressure means more collisions with the walls of the container per minute. This can be a result of more molecules, or molecules traveling faster.

An increase in temperature is an increase in kinetic energy, which means the average speed at which gas molecules travel is faster. Increasing the temperature results in either higher pressure (more collisions with container walls per second) or a larger volume (more distance to travel for each molecule, which keeps rate of collisions with container walls constant).

Nonideal Gases

Using this theory, we can predict the conditions under which gases will not behave ideally.

- 1) Molecules that interact with each other are not ideal. The ideal gas law is only valid if each molecule moves independently. If gas molecules attract or repel each other, the ideal gas law will not hold.
- 2) High pressure. Our approximation that molecules are very far apart is no longer true. This has two consequences. First, molecules that are closer together will have stronger intermolecular interactions. Second, the volume of empty space in which the molecules can travel is reduced as the amount of space occupied by the gas molecules becomes significant.
- 3) Low temperature. Low temperature corresponds to molecules that are moving slowly. Molecules that are moving quickly are more likely to have the energy to overcome intermolecular attractions.

Practice

Umland, p. 181 #31 and 32

Application

Umland, p. 192 #103 and 113

Energy in Chemical Reactions

Objectives: After completion of this material the successful student should be able to

- 1) define enthalpy and entropy, and recognize a spontaneous reaction,
- 2) qualitatively explain the roles of entropy and enthalpy in determining whether reactions are spontaneous,
- 3) analyze the enthalpy and entropy changes in a particular reaction, and
- 4) relate whether a reaction is spontaneous to an analysis of enthalpy and entropy.

Reading assignment: Umland, Chapter 6 pp 196 - 201

Some changes occur spontaneously, such as balls rolling downhill, hot water cooling to room temperature, and combustion of hydrocarbons. Other changes are not spontaneous: balls do not go uphill by themselves, water will not boil unless it is heated, and green plants require energy from the sun to make hydrocarbons like table sugar. Can we predict which processes and reactions will be spontaneous?

What do you already know about spontaneous vs. nonspontaneous reactions?

Potential energy in chemical reactions depends on two things: energy exchanged (usually as heat or light) and disorder. Let's look at these two factors one at a time.

When a chemical reaction occurs, the molecules in the reaction can either gain or lose energy.

It is more common for a reaction to gain or lose heat energy than light energy. Since we can't see heat, how could we tell whether a reaction has gained or lost heat energy?

Chemical Inquiry - Changes in Heat Energy for Spontaneous Reactions

A number of reactants have been prepared for you. Record as many observations as possible, including the initial temperature. Mix the reactants together and record your observations. Write a balanced reaction and determine whether the reactants gained or lost heat energy as they were converted to products. Write your initial and final temperature on the overhead next to the appropriate reaction. When you are finished, compare observations with another lab group and see if you agree with each other's conclusions.

Inquiry Discussion

When the molecules in the reaction transfer heat energy to the surroundings, the temperature of the surroundings increases. When molecules in the reaction absorb heat energy from the surroundings, the temperature of the surroundings decreases.

Exothermic reactions transfer heat energy to the surroundings. Endothermic reactions absorb heat energy from the surroundings.

Most spontaneous reactions are exothermic; however, it is possible for a spontaneous process to be endothermic. This brings us to a discussion of the other determinant of chemical potential energy: disorder.

Systems that are highly disordered are favored in nature, and a sufficient increase in disorder will be spontaneous, even if potential energy is added to the system. Spontaneous change occurs in the direction of increasing probability.

Chemical Inquiry - Increase or Decrease in Enthalpy and Entropy

For our purposes, enthalpy is a change in heat energy. Entropy is the amount of disorder. Decide whether the process or reaction set up for you is spontaneous and decide how the enthalpy and entropy are changing. Be prepared to share a description of your reaction or process and your conclusions with the class.

Inquiry Discussion

The first law of thermodynamics, which we covered in chapter three, states that matter and energy are conserved. The second law of thermodynamics says that spontaneous processes increase entropy of either the system or the surroundings. Spontaneous reactions can be exothermic or increase entropy of the system.

Practice

Umland, pp 196 - 201 # 1-3.

Measuring Enthalpy

Objectives: After completing this material, the successful student will be able to

- 1) use a calorimeter to collect data on enthalpy changes (heats of reaction),
- 2) use specific heats to convert temperature changes to enthalpy changes,
- 3) apply calorimetry to investigating whether enthalpy depends on the reaction sequence.

Reading Assignment: Umland, chapter 6 pp 202-210

For now, we will only deal with enthalpy (changes in heat energy) quantitatively. We will not attempt to measure how much the entropy (disorder) of a system has changed at this point, although we can determine qualitatively whether the entropy has increased or decreased.

In the next chemical inquiry, you will measure how much the enthalpy changes during a reaction. Before you can do that, we need to discuss how different substances absorb heat.

Different substances have differing abilities to absorb heat. Adding one calorie of heat energy to one gram of water raises the temperature of that water by one degree Celsius. Adding the same amount of heat to one gram of iron would raise the temperature of the iron by about two degrees Celsius. Iron and water have different specific heats.

Chemical Inquiry - Calorimetry

Experimental overview: You and your lab partner will measure the enthalpy change per mole of reactants for the acid-base reaction between NaOH and HCl. Some heat given off by the reaction will be dissipated to the surroundings. Decide how you will measure the exact amount of heat lost by your measuring apparatus and write out a detailed set of procedures including chemical hazards and disposal information for all reagents. (Hint: the reading assignment in your text is full of good information.) When your procedures are complete, have them checked by the instructor before proceeding.

Inquiry Discussion:

Exothermic reactions have negative enthalpies because the molecules have less internal energy after the reaction than before (some energy has been given away to the surroundings).

Endothermic reactions have positive enthalpies because the molecules have more energy after the reaction (energy is absorbed from surroundings).

Practice

Umland, pp. 204 - 209 # 5 - 7 and 9 - 12, pp. 224 - 225 # 65 and 67

Application

Umland, pp. 225 - 227 #73, 79, 87, 89, and 91

Calculating Energy Changes

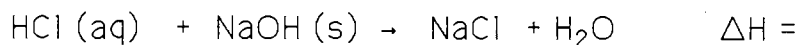
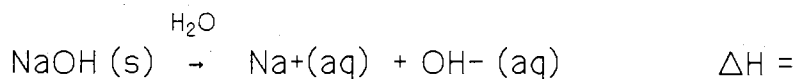
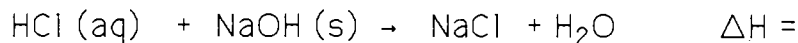
Objective: After completion of this material, the successful student will be able to explain what a state function is and calculate the enthalpy of a reaction from either similar reactions or heats of formation.

Reading Assignment: Umland, Chapter 6, pp 210 -219

Chemical Inquiry - Enthalpy of HCl (aq) and NaOH (aq) vs. NaOH (s)

Your task in this experiment is to determine whether the enthalpy of dissolving NaOH and then reacting it with HCl is the same as reacting solid NaOH with HCl. Using your calorimeter apparatus from the previous inquiry, design an experiment to investigate this question. Have your procedures checked by your instructor before you proceed. Turn in your data before your leave.

Inquiry Discussion



There are two common ways to determine the enthalpy of a reaction. One is by measuring it in a calorimeter, which we have done. The other method involves calculating the enthalpy from constants that you can look up. This is convenient if you do not have the time or the equipment to actually measure the enthalpy. A third, less commonly used method, is to measure the enthalpies of related reactions and then combine them.

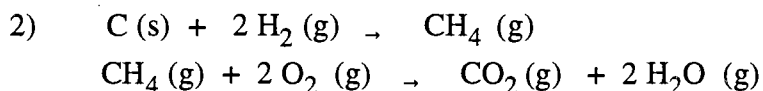
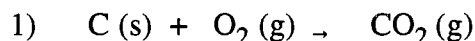
The first law of thermodynamics deals with the conservation of energy (and matter). We cannot have energy suddenly appear or disappear; the total energy at the beginning of a process must equal the total energy at the end of a process.

One of the consequences of this is that the difference in enthalpy (heat of the reaction) between products and reactants is always the same regardless of how the reaction took place.

Forward and reverse reactions require exactly the same amount of energy. The difference is that the energy is added to the system in one direction and given off by the system in the other direction. The enthalpies for reverse reactions are the same, but have opposite signs.

Chemical equations are just like algebraic equations. We can add them together and we can subtract things that are *exactly the same* from both sides.

Which do you think would produce more energy: burning coal directly or converting the coal to natural gas and burning the natural gas? Equations for both options appear below.



One of the things that the law of conservation of energy implies is that the enthalpy difference between individual elements and a compound of those elements is always the same.

If the enthalpy difference between a compound and its component elements is always the same, we only need to measure these enthalpies once, and then they can be recorded in reference books. These enthalpy differences are called enthalpies of formation (ΔH_f).

The enthalpy change for any reaction is always ΔH_f of the products minus ΔH_f of the reactants.

With a neighbor, determine the enthalpy change for the three reactions above, using ΔH_f values from page 214 of the text. Check your answers with another pair when you are finished.

Practice

Umland, pp. 213 - 219 #13 - 17, pp. 224 - 226 # 63, 71, 75 and 85

Application

Umland, pp. 24 - 227 # 59 and 93

You should be able to complete “Stop and Test Yourself” in chapter 6.

States of Matter

Objectives: After completion of this material, the successful student should be able to:

- 1) describe solids, liquids and gases using a kinetic molecular model,
- 2) graph the energy changes necessary to convert one state of matter to another, and
- 3) read a phase diagram.

Reading Assignment: Umland, chapter 12, pp 426 - 430 and pp 444 - 451

Ticket to Class

Draw a picture of a compound as a solid:

a liquid:

a gas:

Definitions

Equilibrium: no net change

Triple point: solid, liquid and gas all exist in equilibrium

Critical point: no longer possible to distinguish between gas and liquid.

Practice

Umland, pp. 445 - 451 # 14 - 18

Application

Umland, pp. 473 - 475 # 91, 119

You should be able to complete #1, 4, 7, 8, and 12 of “Stop and Test Yourself” in chapter 12.

A First Model of the Atom

Objective: Upon completion of this material, the student should

- 1) begin developing a mental model or picture of atoms and molecules, and
- 2) begin developing criteria for evaluating experimental procedures.

Chemical Inquiry - How big is an atom?

Spend the first 10 minutes of the period in groups of two to four. Your group should try to outline an experiment to determine the size of an atom.

Group ideas:

You and your lab partners should spend the next 10 minutes deciding which general experimental strategy you want to use and then write specific procedures to conduct your experiment. Include hazard and disposal information for any chemicals you plan to use, and plan several trials of the same method so that you can get a feel for the precision of your method. When you have recorded your procedures in your lab notebook, and had them approved by your instructor, you may proceed with your experiment.

Inquiry Discussion

Atoms are not really all the same size. They are not hard marbles, either. An atom's size can change depending on its surroundings.

An Atomic Model

Objectives: After completion of this material, the successful student will be able to

- 1) describe the evidence for an atomic model with protons, neutrons and electrons, and
- 2) determine the number of protons, neutrons and electrons in an element.

Reading Assignment: Umland, Chapter 7 pp 230 - 236, Chapter 2 pp 60 -62

Plum Pudding Model

J.J. Thompson discovered that when a large voltage is applied to a piece of metal at one end of an partially evacuated tube, it emits a visible ray of light. When a positively charged plate is placed near the ray, the ray bends toward the plate. Since the rays bend toward the plate, they must carry a negative charge. (1897)

Since metals have a neutral charge, but all can be made to eject electrons if a voltage is applied, it was assumed that there was some positive portion of the molecule that electrons were embedded in.

Gold Foil Experiment

Rutherford had several graduate students perform an experiment which seemed to suggest a small positively charged nucleus surrounded by a large region of electrons. (1909)

Electron Mass

Millikan measured the charge and mass of the electron by closely observing oil drops. He measured the mass of an electron as 5×10^{-4} atomic mass units. All electrons have the same negative charge, so we say the charge of an electron is -1. (1913)

The mass and charge of a proton were measured by a similar method. Masses for most elements were much too large to be accounted for by protons and electrons, so it was assumed that another particle, with no charge, existed in the nucleus with the protons.

Information from the Periodic Table

The periodic table tells us how many protons, neutrons and electrons are in an atom of each element. The number above the chemical symbol is the number of protons. The identity of an atom is determined by the number of protons it contains.

The number underneath the chemical symbol is the relative average atomic mass. It is an average because not all atoms of one element have the same mass. Two atoms of the same element have different masses if they have different numbers of neutrons. Protons and neutrons have approximately the same mass; each weighs about one amu.

Draw and label a picture of an atom incorporating as much of the information we have acquired so far as possible. When you are finished, compare your representation with your neighbors. Discuss any differences in your pictures.

Practice

Umland, chapter 2, p 62 # 31 and 32; p 72 # 11 and 12

Application

Umland, p 262 #61, p 264 #75

Planetary Model of the Atom

Objectives: After completion of this material, the successful student should be able to

- 1) trace the development of evidence that led Bohr to adopt a planetary model of the atom, and
- 2) relate the wavelength of a photon to its energy.

Reading Assignment: Umland, Chapter 7 pp 236 - 249

Chemical Inquiry - Atomic Spectra

A number of spectrometers are set up. Record the element and your observations.

Q: How do they make fireworks of different colors?

Definitions and Background Information

Visible light is part of the electromagnetic spectrum. Radio waves, micro waves and infrared are all types of electromagnetic radiation with less energy than light. Ultraviolet and x-rays have more energy than visible light. All electromagnetic radiation travels in waves, at the speed of light.

Wavelength is the amount of distance between the crest of one wave and the crest of the next wave.

Frequency is how many waves pass a given point in a second.

Wavelength times frequency equals the speed of light.

$$\lambda \nu = c$$

meters(1/s) = m/s

Higher energy electromagnetic radiation has a higher frequency and shorter wavelength.

Photoelectric Effect

Electrons are emitted from a piece of metal when light strikes the metal. Not just any light causes this photoelectric effect. For instance, ultraviolet light is necessary to eject electrons from aluminum. The more intense the UV light is, the more electrons are emitted. But no matter how bright the light, visible light will never knock electrons out of aluminum.

Photons

Einstein suggested that these experimental results were not so strange if we think of light as tiny particles or packages of energy, with a collection of these packages traveling in waves. One package of light then, must have a certain minimum energy in order to cause an electron to be ejected from the metal.

Planck related the energy of one package of light to its frequency.

$$E = h\nu$$

The constant h in this equation is called Planck's constant and it equals 6.63×10^{-34} Joules (seconds).

Bohr's Model

As Bohr considered the structure of the atom in 1913. He knew that

- 1) the nucleus is small, dense and positively charged,
- 2) negatively charged electrons take up the space around the electron,
- 3) in order to fill the space around the nucleus, the electrons are probably moving rapidly, and
- 4) only certain wavelengths of light are absorbed or emitted by a given atom.

The biggest puzzle to Bohr was what kept the negatively charged electrons from falling into the positively charged nucleus. He guessed that this puzzle was related to the fact that only certain wavelengths of light were absorbed or emitted by an atom.

He suggested that electrons might be confined to certain paths or orbits around the nucleus. As an electron moved from one defined orbit to another, a specific amount of energy would be absorbed or emitted.

Practice

Umland, p. 262 - 265 #57, 85, 93, 99

Application

Umland, p. 264 - 265 # 67, 77,

Quantum Mechanical Model of the Atom

Objectives: After completing this material, the successful student will be able to

- 1) articulate the Heisenberg Uncertainty Principle and use it to describe the location of an electron,
- 2) describe what is meant by wave-particle duality,
- 3) identify at least one problem with the Bohr model of the atom,
- 4) explain what an electron orbital is and describe the shapes of s, p and d orbitals,
- 5) use the periodic table to write electron configurations,
- 6) determine whether an atom is paramagnetic or diamagnetic, and
- 7) describe the atom using the quantum mechanical model.

Reading Assignment: Umland, Chapter 7 pp 250 - 259, Chapter 8 pp 269 - 276, excerpts on reserve in the library.

Ticket to Class

Describe what a "thought experiment" is in your own words. Summarize the thought experiments described in the supplemental reading.

Wave-Particle Duality

In 1927, Heisenberg formulated his Uncertainty Principle, which states that we can never know the exact position and velocity of a particle. The uncertainty of position and the uncertainty of velocity must always be greater than or equal to Planck's constant.

In the same year, 1927, a stream of electrons was diffracted. Diffraction is a property of waves, so in this case a particle with a measurable mass, the electron, is behaving like a wave.

One way of interpreting these observations is that everything, matter and energy, is both a wave and a particle. What we see, wave behavior or particle behavior, depends on what we look for.

Schroedinger's Model

Bohr's model of the atom is not perfect. First, it only predicts the spectral lines for a few elements. A larger problem, however, is that it tries to assign definite distances from the nucleus to the allowed electron orbits.

Schroedinger developed equations that represented the electrons within an atom as standing waves. Solutions to the Schroedinger equation tell us the probability of finding an electron in a small region of space.

The probability of finding an electron is highest close to the nucleus and decreases as we move further away from the nucleus. The probability does not become zero until we move infinitely far away from the nucleus.

Since it is inconvenient to talk and think about infinity, we have arbitrarily chosen the volume around the nucleus where an electron will be found 90% of the time to describe the location of an electron. This space around the nucleus is called an orbital. Each orbital can contain a maximum of two electrons.

Four variables called quantum numbers are used to describe orbitals:

1) n describes orbital size. Although we cannot determine the exact position of an electron, we can determine an average distance from the nucleus based on probabilities. An $n=2$ electron, in the second shell, is farther from the nucleus than an $n=1$ electron, in the first shell. Electrons that are, on the average, farther away from the nucleus have higher energies.

2) l describes the orbital's shape. s , p and d orbitals have different shapes.

3) m_l describes the orientation of the orbital in space. Spherical s orbitals, for instance, can only have one orientation. Therefore, there is only one s orbital in each shell. p orbitals have three possible orientations and there are three p orbitals in each shell.

4) m_s describes the spin of an electron, which is either $+\frac{1}{2}$ or $-\frac{1}{2}$. Electrons must have a different spin to occupy the same orbital.

Electron Configurations

Sometimes it is useful to write out how electrons around an atom are arranged. By looking at the periodic table, we see the first shell holds two electrons (there is only a 1s orbital). The second shell holds a total of eight electrons, two in a 2s orbital and six in three 2p orbitals. We write the electron configuration of lithium, with three electrons as $1s^2 2s^1$. Neon, with 10 electrons, is written $1s^2 2s^2 2p^6$.

Sometimes, not all the orbitals are full. Lithium, for instance, has an unpaired electron in the 2s orbital. Atoms with unpaired electrons are paramagnetic, which means that they respond to a magnetic field.

In an atom of oxygen, there are a total of eight electrons. Two go in the first shell, two in the s orbital of the second shell, and the remaining four must be placed in the three p orbitals. Since electrons repel each other, electrons in the same orbital must have the opposite spin and each orbital should get one electron before any orbital has two.

O: 6 electrons



Choose an element from the first two rows of the periodic table. On a piece of paper, draw and label one or more pictures which represent the best mental model you can form of an atom of that element. Write the electron configuration for that element. When you are finished, compare and discuss your model with those of your neighbors.

Practice

Umland, p. 259 # 18; p. 270 - 272 #1, 5 - 7 and 9; p. 300 # 71 and 79

Application

Umland, p. 300 # 75 and 77; p. 302 # 83 and 95

You should be able to complete “Stop and Test Yourself” in chapter 7.

Trends in the Periodic Table

Objectives: After completion of this material, the successful student will be able to

- 1) describe properties of active metals, transition metals and nonmetals,
- 2) identify how several properties of elements change within the periodic table,
- 3) explain the concept of effective nuclear charge, and
- 4) use the periodic table to compare ionic and atomic radii.

Chemical Inquiry - Periodic Properties

For each of the following elements: sodium, potassium, magnesium, calcium, iron, copper, aluminum, phosphorus, sulfur, and iodine, place a small (pea size) chunk in water and observe any results. Use pH paper to test whether the resulting solution is acidic or basic.

Choose another property you want to investigate. Your investigation should include enough elements or compounds that a trend across the periodic table or within a column in the periodic table will be evident. You will have a list of available chemicals, but you must look up hazard and disposal information for any reagents you select. Have your instructor approve your procedures before you begin. Collect your data and observations carefully, and be prepared to present your findings to the class.

Inquiry Discussion:

The behavior of elements is most strongly determined by the number of electrons in the outermost shell. The modern periodic table is arranged so that elements with the same valence electron configuration are in the same column of the periodic table.

Trends across the periodic table are dominated by the effective nuclear charge experienced by outer shell electrons.

Low effective nuclear charges, farther to the left on the periodic table, mean electrons that are farther away from the nucleus. This results in large atoms that lose electrons easily. High effective nuclear charges result in smaller atoms that tend to gain electrons to form negative ions.

Trends down the periodic table depend on the increasing distance of outer shell electrons from the nucleus. Atoms are larger as you move down the periodic table.

Positive ions have lost electrons and are smaller than the neutral element. Negative ions are larger.

Application

Umland, p. 302 - 303 #85, 87 and 105

You should be able to complete “Stop and Test Yourself” in chapter 8.

Chemical Bonding

Objectives: After completing this material the successful student should be able to

- 1) define and identify ionic and covalent bonds,
- 2) use Lewis structures to predict bonds and the existence of unshared pairs, and
- 3) draw appropriate resonance structures and explain their physical significance.

Reading Assignment: Umland, chapter 9 pp 308- 310, 314 - 317, 321 - 324 and 328 - 334

Ticket to Class

We have drawn a model of an individual atom, but most chemistry is about compounds.

What would a picture of a compound look like? How would it be similar to a picture of an individual atom and how would it be different? In the space below, summarize your ideas on these questions, in writing, and be prepared to discuss your summary with a classmate.

In a compound, the atoms are attached to each other. There are two common ways to describe how this happens:

- 1) Two ions of opposite charge can associate with each other. The attraction of unlike charges holds the atoms close together. This is called an ionic bond between atoms.
- 2) Electrons from two atoms can be “shared.” In this case, the negative electrons from both atoms are under the influence of both nuclei. This is called a covalent or molecular bond between atoms.

Ionic bonding is more likely between ions with opposite charge (a metal and a nonmetal), usually in a solid. Covalent bonding occurs whenever there are no charged atoms present. Covalent compounds can be solids, liquids, or gases.

Ionic Bonds

In ionic compounds, ions with opposite charges are held together in large three dimensional arrangements. Each atom keeps its own electrons, presumably in orbitals that are similar to the orbitals for a single ion of that element.

Covalent Bonds

One way to think of covalent bonding is that the shared electrons fill up available orbitals for both atoms. Let's think about a molecule of oxygen as an example.

A molecule of oxygen contains two atoms. Both oxygen atoms, by themselves are "missing" two p-orbital electrons. If both atoms were to share two of their p electrons with the other atom, then both atoms would have a full outer shell. We expect a full shell to be fairly stable, since the noble gases, which also have full shells, are so unreactive.

Thinking about atoms sharing electrons to fill the available orbitals of both atoms gives us a handy way to predict how many covalent bonds we can expect between two atoms. Multiple bonds indicate that atoms are held closer together. Each two electrons shared counts as one bond, so oxygen with four shared electrons has a double bond between the two atoms.

Most bonds between atoms are neither strictly ionic nor strictly covalent. Whether they are more like ionic bonds or covalent bonds is determined by the difference in the atoms' electronegativity. Your text has a chart of electronegativity on p. 318.

Predicting Number of Bonds Between Atoms

Lewis structures are a set of rules which help you predict the number of bonds between atoms without writing out the electron configuration of each atom in the compound.

To draw a Lewis structure:

- 1) Write down all the atoms so that they can be properly linked together. If there is a unique atom (like sulfur in SO_2), that is usually your central atom. The less electronegative atom will

usually be in the center of the Lewis structure.

- 2) Add all the valence electrons from all atoms together.
- 3) Put two electrons between each two linked atoms.
- 4) Add electrons to the outer atoms until there are eight around each atom or until you use up the number of electrons you determined in step 2. Electrons between atoms are counted as part of the total eight for both atoms.
- 5) If you still have electrons left over, add electrons to the central atom until you run out. Sometimes the central atom can have more than eight electrons if it is in the third row of the periodic table or lower.
- 6) If any of the atoms in the molecule do not have eight electrons, move additional pairs of electrons between atoms until you can count eight around all atoms.
- 7) Decide whether resonance structures are possible.

Practice

Umland, p. 310 # 4 - 5, p. 315 - 316 # 7 - 11, p. 324 - 325 # 16 - 18, p. 331 - 334 # 24 - 28, p. 349 # 81 and 89

Application

Umland, p. 349 # 99 and 103

Molecular Models

Objectives: After completing this material, the successful student should be able to:

- 1) begin building a personal model of molecular structure which is expected to change as new information is gathered, and
- 2) describe the VSEPR theory and use it to predict the shapes of molecules.

Reading Assignment: Umland, chapter 9, pp 334 - 341

Ticket to Class

Evaluate your model of a molecule from the previous page based on the information about bonding between atoms discussed so far. Record any alterations or corrections to your model in the space below. Choose a compound, draw the Lewis structure and describe that molecule as completely as possible.

VSEPR (Valence Shell Electron Pair Repulsion) theory simply notes that electrons repel each other, and that things that repel each other move as far away from each other in space as possible.

Chemical Inquiry - Molecular Shapes

Methane (CH_4) has four hydrogen atoms around a central carbon atom. Draw a Lewis structure for this compound. Use foam balls and toothpicks to compare possible three dimensional arrangements of the atoms. Decide which is the most likely arrangement, remembering that electrons repel each other. Be prepared to share your ideas with the class.

Now draw Lewis structures and determine the three dimensional shapes of carbon dioxide (CO_2), water, ammonia, and sulfur trioxide (SO_3). When you are finished, discuss your conclusions with someone nearby.

Inquiry Discussion

In order to predict the shape of a molecule:

- 1) Draw an accurate Lewis structure.
- 2) Arrange atoms so that electron pairs (bonding and unshared) are as far away from each other as possible. Note: Double and triple bond pairs do not move away from each other; they are held close together between the atom.
- 3) Determine the overall shape of the atoms (not the electrons) and approximate bond angles.

Practice

Umland, p. 340 # 29-31 and 33

Polar and Nonpolar Molecules

Objectives: After completion of this material, the successful student should be able to:

- 1) define what is meant by polar and nonpolar bonds,
- 2) use electronegativities to place partial positive and negative charges, and
- 3) determine whether a molecule is polar based on molecular shape and electronegativities.

Reading Assignment: Umland, chapter 9 pp 317 - 320 and 341 - 345

Atoms do not share electrons equally and electrons that are shared are distributed around the molecule in molecular orbitals, which are not the same shape or size as the orbitals of the two atoms. Fortunately, we can understand a lot of chemistry without discussing the exact shape of molecular orbitals. Even at an introductory level, however, we need to explore unequal sharing of electrons.

In a water molecule, for example, electrons are shared between oxygen and hydrogen. However, the electrons spend more time close to the oxygen. Oxygen, with the electrons close to it more than half the time, ends up with a slight negative charge. The hydrogens have a slight positive charge.

A bond between two atoms with similar electronegativities will be *nonpolar*, which means that electrons are shared equally, with no partial charges. A bond between atoms with very different electronegativities will be very *polar*, which means that the partial positive and negative charges will be relatively large. The larger electronegativity has the partial negative charge.

In order for a molecule to be polar, it must contain polar bonds that do not cancel each other.

Practice

Umland, pp. 319 - 320 # 12 and 13; p. 345 # 34

Application

Umland, p. 351 # 113

You should be able to complete #1-12 of "Stop and Test Yourself" in ch. 9.

Attraction Between Molecules

Objectives: After completion of this material, the successful student should be able to:

- 1) identify the types of intermolecular attractions between given molecules, and
- 2) describe dynamic equilibrium.

Reading Assignment: Umland, chapter 12, pp 430 - 444

Chemical Inquiry - Boiling Point

Measure the boiling points of two liquids. Use the CRC to look up eight more.

List any patterns you see in the data. Be prepared to share these with the class. Develop a theory or "explanation" for any of the patterns you see.

Intermolecular Forces

Matter is generally sticky. At low temperatures where molecules move slowly, they tend to clump together rather than drifting far apart. We explain this clumping tendency as electrostatic attraction between positive and negative “ends” of molecules.

Draw a diagram of water, and label any partial charges. Now draw another water molecule next to the first one, as you would expect it to look in a glass of water. Compare your results with a neighbor.

Now draw sodium chloride dissolved in water.

Dipole - Dipole and Ion - Dipole Attractions

Opposite charges attract; like charges repel. The attraction between opposite (full or partial) charges holds ions and/or polar molecules together.

Hydrogen bonding

Water, DNA, proteins, alcohols, and lots of other important molecules can hydrogen bond. Hydrogen bonding is just a special name for an extremely strong dipole-dipole interaction.

Attraction Between Nonpolar Molecules

Nonpolar molecules also stick together because of attraction of opposite charges. Polar molecules have permanent dipoles. Nonpolar molecules have induced (temporary, changing) dipoles.

The attraction between nonpolar molecules is often weaker than the attraction between polar molecules. Larger molecules attract each other better than small ones because there is more surface area for contact between the two electron clouds, and therefore better temporary dipole formation.

Viscosity and Diffusion

Strong IMF = high viscosity. The shape of a molecule is also important to viscosity. Big, flexible molecules twist around each other, slowing the movement of both. Small, round molecules slide past each other more easily. Diffusion is slower in more viscous solutions.

Evaporation and Condensation

How fast a liquid evaporates depends on the intermolecular forces between the molecules. Strong IMF = slow evaporation (and high boiling point).

Dynamic Equilibrium

When the rates of two opposite processes (like evaporation and condensation) are equal, we have dynamic equilibrium.

DYNAMIC - there is plenty of action; the system is not "resting"

EQUILIBRIUM - there is no net change

Practice

Umland, pp. 433 - 444, # 1-13

Application

Umland, p. 473 - 475 # 89, 97, 107 and 121

You should be able to complete # 2, 3, 5, 6, and 11 of "Stop and Test Yourself" in chapter 12.

Properties of Solutions

Objectives: After completion of this material, the successful student should be able to:

- 1) predict and rationalize solubilities by comparing strength of intermolecular forces for the solvent and solute,
- 2) predict the effect of temperature and pressure on solubility,
- 3) use a variety of units to describe concentration, including ppm and molality, and
- 4) define, identify and perform calculations for colligative properties.

Reading Assignment: Umland, chapter 13, pp 478 - 493 and 500 -508

Ticket to Class

Choose a liquid and a solid or two liquids that you know are soluble. Draw Lewis structures for both molecules, and determine the molecular geometry and distribution of partial charges.

Draw a picture of several molecules of each substance in separate containers. Determine what type of intermolecular forces are involved.

Draw a picture of several molecules of each substance together in the same container.
Determine what type of intermolecular forces are involved.

Predicting Solubility

In order for a solution to form, three things need to happen:

- 1) The molecules of the solvent must separate.
- 2) The molecules of the solute must separate.
- 3) Molecules of the solvent and solute need to move close together.

In terms of energy, these three steps can be summarized as

- 1) increase in potential energy
- 2) increase in potential energy
- 3) decrease in potential energy

Two substances are soluble if the decrease in potential energy in the third step is larger than the increases from the first two steps.

This is why “like dissolves like.” In molecules with similar IMF, the energy difference between the separate compounds and the solutions is likely to be small. The increase in entropy will outweigh a small unfavorable energy difference.

Temperature and Pressure

Solids and liquids are more soluble at higher temperatures. Pressure does not affect the solubility of solids or liquids.

Gases are less soluble at higher temperatures. Higher pressure increases the solubility of a gas.

Units of Concentration

$$\text{molarity} = \frac{\text{moles solute}}{\text{liter solution}}$$

$$\text{molality} = \frac{\text{moles solute}}{\text{kg solvent}}$$

$$\text{mole fraction} = \frac{\text{moles solute}}{\text{total moles in solution}}$$

Colligative Properties

Colligative properties depend only on the NUMBER of particles, not on their identity. Each ion or molecule counts as a particle.

Vapor Pressure Lowering

Solutions have a lower vapor pressure than the pure solvent.

$$\text{vapor pressure of solvent in solution} = \text{mole fraction solvent (vapor pressure of pure solvent)}$$

Boiling Point Elevation/Freezing Point Depression

Adding a solute raises the bp of the solvent and lowers the freezing point.

$$\Delta \text{ boiling point} = K_{bp}m \quad \Delta \text{ freezing point} = K_{fp}m$$

Osmotic Pressure

Unless there is the same concentration of particles on both sides of a semipermeable membrane, solvent will diffuse from the less concentrated solution to the more concentrated solution.

Practice

Umland, p 484 - 488 # 1 - 5, p. 490 # 7 -9 and 11, p. 493 - 494 # 13 - 15, p. 501 - 509 # 21 - 29

Application

Umland, p. 519 # 81, 103, 105, 117 and 119

You should be able to complete # 1 - 13 of "Stop and Test Yourself" in Chapter 13.

Reactions Which Transfer Electrons

Objectives: After completion of this material, the successful student should be able to:

- 1) assign oxidation numbers,
- 2) define and identify oxidation and reduction,
- 3) balance oxidation-reduction equations, and
- 4) perform stoichiometric calculations for oxidation-reduction reactions, including redox titrations.

Reading Assignment: Umland, chapter 11, pp 393 - 408

Oxidation reduction reactions are crucial for human life. The primary mechanism our cells use to transform the food we eat into usable energy is a series of oxidation-reduction reactions called the electron transport chain.

Chemical Inquiry - Redox Reactions

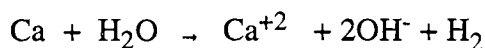
A number of solutions and solid reagents are available. With a partner, place the solid reagents in each solution and observe what happens in each case.

After you have finished recording your observations, discuss your results with another pair. Try to determine what, if any, patterns you see in the data.

Inquiry Discussion

Oxidation means a loss of electrons; reduction is a gain of electrons. Since electrons cannot appear out of nowhere or vanish into thin air, oxidation and reduction always occur together.

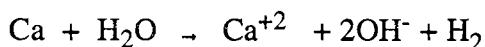
In a redox reaction, electrons are transferred and charges on ions change. Let's look at dissolving calcium in water as an example.



The total charge on both sides of the reaction is balanced. Calcium does not have the same charge on both sides; calcium has lost two electrons.

The electrons went somewhere. Since water is the only reactant, water must have accepted the two electrons that calcium donated.

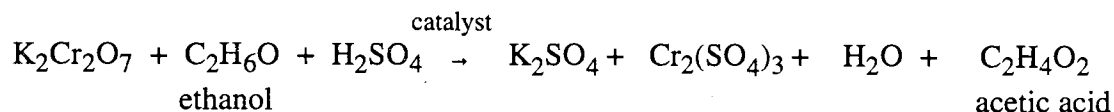
The process of assigning oxidation numbers following the rules on p. 395 in your text will lead us to the same conclusion.



Whether we say that hydrogen or oxygen "keeps" the electrons from calcium is obviously just a bookkeeping system.

Look at Number 11.33 on p. 419 of your text. Using oxidation numbers, determine which are redox reactions and which are not.

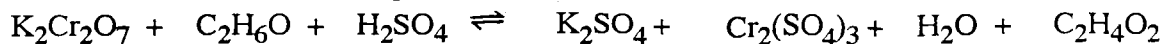
Redox reactions can be difficult to balance by inspection. Give this one a try:



Without a balanced equation, we cannot do quantitative stoichiometric calculations. In this case that would be a disaster, because this is the reaction used in a Breathalyzer unit. (See p.417 of your text for more details.)

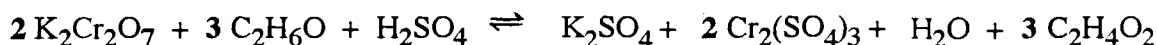
Using oxidation numbers is usually an easy way to balance redox reactions.

1) Assign oxidation numbers to all atoms.



2) Identify what is oxidized and what is reduced. Balance those atoms.

3) Find the increase and decrease in charge. Change coefficients so that the charge increase equals the charge decrease.



4) Finish balancing by inspection. DO NOT change the coefficients of the redox reagents. Balance hydrogen and oxygen last.

You can add water to either side of any aqueous solution to help balance H and O. You can also add H^+ to acidic solutions or OH^- to basic solutions. (How do you know whether a reaction is in acidic or basic solution?)

5) Check that the charge and number of atoms are the same on both sides of the equation.

Once your equation is balanced, all stoichiometric calculations are the same as they would be for acid-base or precipitation problems.

Practice

Umland, pp. 397 - 408 #1 - 12; p. 417 #21

Application

Umland, p. 421 # 56

You should be able to complete #1-4 and 8-14 of "Stop and Test Yourself" in chapter 11.

2
VITA

Carolyn Herman

Candidate for the Degree of

Doctor of Education

Thesis: REFORM PRACTICES IN CHEMISTRY EDUCATION IN
INTRODUCTORY COLLEGE AND UNIVERSITY COURSES

Major Field: Higher Education

Biographical:

Education: Received Bachelor of Arts in biology from the University of Missouri -
Columbia in May 1986; received Master of Science in biochemistry from Colorado
State University in December 1988; completed requirements for Doctor of
Education at Oklahoma State University in July 1996.

Experience: Oncley Professor of Chemistry at Southwestern College in Kansas;
Director of an undergraduate research project funded by the Forest Products
division of the USDA; Participant in Alliance for Science in Kansas, a consortium
of independent colleges which sponsors Project ChemKits, a program to supply
instrumentation to area high schools for short-term use.